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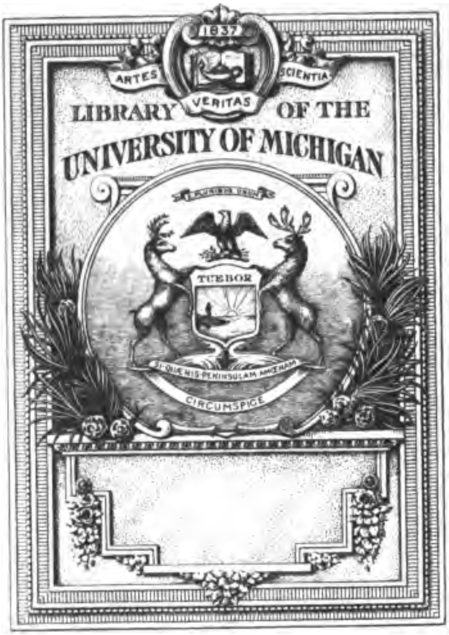
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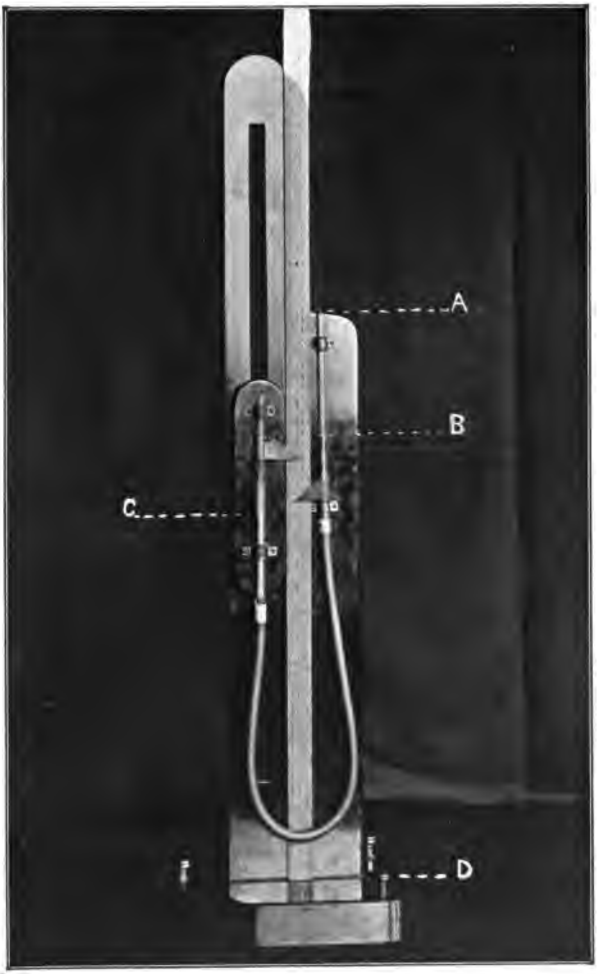

FIG. 54

Frontispiece

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PART I

PHYSICS

ELEMENTARY SCIENCE FOR PUPIL TEACHERS

PART I

PHYSICS

MEASUREMENT AND MATTER

CHAPTER I

MEASUREMENT OF LENGTH

Measurement.—Accurate measurement plays a most important part in the progress of all science; indeed, most of the great discoveries of science have been brought about by its aid. In a book of this kind it is scarcely possible to lay too much stress on the necessity for careful and exact measurement, since the majority of the experiments described require for their performance the measurement of some quantity or other.

Distance or length is the quantity whose measurement will be dealt with first. It is measured by comparison with a fixed length called the *unit of length.* Thus we might say that the length of a bench is twenty times the length of a certain pencil, meaning that twenty exactly similar pencils placed end to end in a straight line would stretch from one end of the bench to the other. Here the length of the bench is measured by comparing it with the length of a certain pencil, the length of the particular pencil being taken as the unit of length. But the length of the bench might be twenty-five times the length of another pencil, or thirty-six times the length of a stick of chalk; so that the expression of a length depends upon the unit used.

Therefore, for general use and to avoid confusion, it is evident that some definite unit of length must be chosen. Such a unit is called a *standard unit*.

There are two recognised systems of units in use in this country, the **British system** and the **Metric system.**

The British unit of length is called the **Yard,** and is the distance between two fine scratches on gold plugs let into a bronze bar when its temperature is 62° F. The bar is kept in the Standards Office of the Board of Trade at Westminster.

One-third of this distance is called a foot, and an inch is one-twelfth of a foot.

1 yard = 3 feet = 36 inches.
1,760 yards = 1 mile.

In the Metric system the standard unit of length is the **Metre,**[1] which was originally intended to be the 10,000,000th part of the distance from the earth's equator to the North Pole, measured along the meridian of Paris; that is, the 10,000,000th part of the quadrant of the earth represented by the arc EPN in Fig. 1.

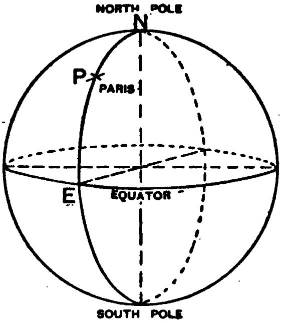

FIG. 1

A bar of platinum was carefully made to be of this length at the temperature of melting ice, and was deposited in the French Archives. The length of this bar is the **standard metre.**

We now know that a quadrant of the earth is more than 10,000,000 times the length of this bar, for it has since been found that a slight error was made in calculating the length of the quadrant. Thus the standard metre is not what it was originally intended to be, namely, a natural unit connected with the size of the earth. It is, to all intents and purposes, like the British yard, an artificial or arbitrary unit.

The metre is divided into 10 equal parts, each being called a *decimetre* (Latin, *decem*, ten); a decimetre is divided into 10 equal parts, each called a *centimetre* (Lat., *centum*, a hundred); each centimetre is also divided into 10 equal divisions, each called a *millimetre* (Lat., *mille*, a thousand).

---

[1] Greek, *metron*, a measure. Pronounced "meeter" in this country.

Lengths which are 10, 100, and 1,000 times the metre are also used, and are given names with Greek prefixes.

The metric table of lengths is as follows:—

10 millimetres (mm.) = 1 centimetre (cm.)[1]
10 centimetres = 1 decimetre (dm.)
10 decimetres = 1 metre (m.)
10 metres = 1 dekametre (Dm.)
10 dekametres = 1 hectometre (Hm.)
10 hectometres = 1 kilometre (Km.)

We thus see that all the multiples and sub-multiples of the metre are given by powers of ten; for this reason the Metric system is called a *decimal* system, and in this its advantage lies, and not in the actual length of the unit; and, as will be seen later, the metre has other measures and weights derived from it by the most simple methods of multiplication and division.

The inconvenience of the British system has brought about in this country the use of the Metric system for all scientific measurements. In fact, the Metric system is universally employed for purely scientific investigations. It is also in use in many countries for ordinary purposes.

In physical determinations the metre is too large a unit, so that lengths are usually expressed in centimetres.

Expt. 1.—You are provided with a boxwood metre scale divided into decimetres, centimetres and millimetres along one edge, and having inches divided into tenths marked off along the other edge (Fig. 2).

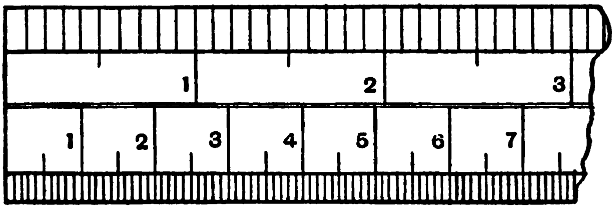

FIG. 2

Examine the scale, and note that every centimetre division is numbered; also that the metre is longer than the yard. Draw straight lines

---

[1] Abbreviations commonly used.

1 inch and 1 cm. in length, and note that the inch is longer than the centimetre.

Also draw a straight line 1 dm. 1 cm. 1 mm. in length.

To illustrate the advantage of the decimal system of subdivision, suppose we are asked to write down the length of the blackboard. On careful measurement we find that it is 2 metres, 3 decimetres, 5 centimetres, 4 millimetres. We should not write it

m.	dm.	cm.	mm.
2	3	5	4

but as 2354 mm., or 235·4 cm., or 23·54 dm., or 2·354 m.

Thus in reducing a length the figures remain the same, and it is only necessary to move the decimal point.

Expt. 2.—Write down the following lengths (1) in metres, (2) in decimetres, (3) in centimetres, (4) in millimetres.

	Km.	Hm.	Dm.	M.	dm.	cm.	mm.
(*a*)	4	3	8	7	0	2	6
(*b*)	0	0	0	6	5	0	3
(*c*)	0	0	0	0	3	6	7

Expt. 3.—By examination of the scale find the length of 1 m. in inches and decimal of an inch; also the length of 1 inch in cm. and decimal.

How to use a scale to measure directly the distance between two points.—In the first place do *not* make use of the end division, which is usually inaccurate, even when the scale is new, for the ends are seldom cut quite square; moreover, constant use wears away the edges of the scale, particularly at the ends. Therefore it is better to measure from a centimetre division.

Secondly, suppose we wish to determine the distance between two points A and B, the scale should not be used as shown in Fig. 3; for, unless the eye be placed vertically above the points, the correct readings will not be obtained. In this case the error is due to the thickness of the scale. In the figure the reading for the point A, when the eye is correctly placed, as at E, is 1 cm. If, how-

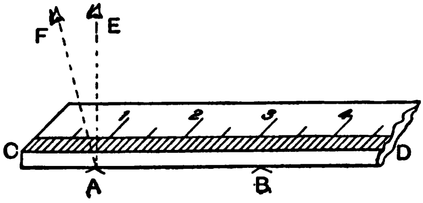

FIG. 3

ever, the reading be taken from a position **F**, slightly out of the vertical, a different reading, 0·93 cm., is the result.

If the edge **CD** of the scale be placed in contact with the points **A** and **B** (Fig. 4), this error cannot arise.

Notice also in this figure that the length of the straight line joining **A** and **B** is more than 21 mm. and less than 22 mm. With practice it is easy to imagine the last millimetre divided into ten equal parts, and thus to estimate that the length of **AB** is 21 mm. and 4 of these equal parts, or 21·4 mm. In this way it is possible to estimate to hundredths of a centimetre.

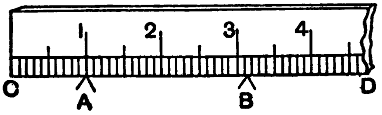

FIG. 4

Expt. 4.—Measure the length of this page in four places. Suppose the results obtained are 17·92 cm., 17·93 cm., 17·92 cm., and 17·90 cm.

All the determinations do not agree, so that we cannot say that the length of the page is 17·92 cm., nor yet 17·90 cm. In such a case we take the *average or mean* of all the measurements, thus:—

Sum of the four measurements = 71·67 cm.

Average or mean ,, ,, ,, $= \dfrac{71{\cdot}67}{4}$

$= 17{\cdot}9175$ cm.

Hence, average length of page = 17·92 cm.

Notice that $\dfrac{71{\cdot}67}{4} = 17{\cdot}9175$, and that the mean value is only taken to the second place of decimals. This is because our measurements were taken to the nearest hundredth of a centimetre. A result of 17·917 cm. would imply that we had measured to the thousandth part of a centimetre.

Also the mean is taken as 17·92 and not as 17·91, for the reason that 17·917 is nearer 17·92 than 17·91.

Usually, in the measurement of any quantity, a number of observations are made, and the mean value found, as above.

Also notice that the fourth measurement is put down as 17·90 cm. and not as 17·9 cm., for they do not mean quite the same thing. *The former shows that an attempt has been made to estimate to the hundredth part of a centimetre. The latter reading means that no such attempt was made and the hundredths of a centimetre were not taken into account.* The reading 17·90 cm. means that the hundredths were sought for, and none were found.

Measure the length and width of your note-book in three places, and calculate the mean length and width.

Expt. 5.—Draw in your note-book three straight lines **AB**, **CD**, **EF**. Using your small boxwood rule divided as in Fig. 2, measure carefully the length of each line in inches and decimal of an inch, also in centimetres and decimal.

The length of a line in cm. divided by its length in inches evidently gives the number of cm. in one inch. Calculate this number in each case, and enter your results in a table drawn up as follows:—

Line measured.	Length in inches.	Length in cm.	Number of cm. in 1 inch.
AB CD EF			

Expt. 6.—Measure in cm. and in inches

(1) The distance between the 1st and 6th lines in your note-book.
(2) „ „ 1st and 8th „ „ „
(3) „ „ 1st and 10th „ „ „

From these measurements find the number of cm. in 1 inch. Make a suitable table and insert your results.

Expt. 7.—Estimate the length and breadth of (1) the blackboard, (2) the room, (3) the top of your bench. Also estimate the lengths of straight lines drawn on the blackboard. Then make the measurements with a metre scale, and see if your estimations are correct.

ADDITIONAL EXPERIMENTS

8. Without measuring from a scale, draw a line 1 inch long and another 1 cm. in length. Then apply your scale and see whether you have estimated the lengths correctly.

9. Measure, in four or five places, the thickness of the pages of this book, and hence obtain a mean value. Find the number of pages, and then calculate the thickness of one page.

10. Estimate and then measure (*a*) the diameter and thickness of a halfpenny, (*b*) the length and width of the laboratory.

11. Mark four dots, **A**, **B**, **C**, **D**, on your note-book. Then measure the distances of **A**, **B**, and **C** from **D**, in cm. and decimal.

EXAMPLES I

1. Express 1 mm. as a decimal of (1) a metre, (2) a kilometre.

2. How many decimetres are there in 0·0865 dekametre?

3. Reduce 6·75 metres to mm.

4. Express 1 cm. as a decimal of a dekametre.

5. The length of a football ground is 1 hectometre. Write down the length in mm.

6. The diameter of a halfpenny is 1 inch. How many such coins placed in a row would stretch 5·08 decimetres?

7. A person walks at the rate of 4 miles per hour. How many kilometres per hour is this?

8. Convert 1 kilometre to yards, feet, and inches.

9. A line is 4·7 inches long. What is its length in cm. and decimal?

10. If the length of the playground is 60·3 metres, what is the length in feet?

11. If the height of a column of mercury is 760 mm., what is it in inches?

12. Express 1 mm. as a decimal of (1) a dekametre, (2) a hectometre.

Expt. 12.—To measure the diameter of a sphere.

Using only a scale, it is impossible to measure with any degree of accuracy the diameter of a sphere. By the following method, however, it is easily obtained. A and B (Fig. 5) represent two blocks, the ends of which are accurately squared, *i.e.* the face C is exactly at right angles to A, and D is exactly at right angles to B. The blocks are arranged to touch the graduated edge of the scale with the sphere in contact with the ends C and D.

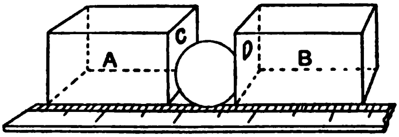

FIG. 5

The diameter is then the distance between these ends, and may be found by reading the scale.

It is to be noted that the height of each block must be greater than the *radius* of the sphere.

Expt. 13.—Measure the diameter of a cylinder by the above method.

Measurement of the lengths of curved lines.—A curved line may be supposed to consist of a large number of very short straight lines. Thus we can imagine the curve

AXB (Fig. 6) to be made up of short lines CD, DE, EF, etc., and the sum of their lengths will be nearly equal to the length of the curve.

If the lines be made shorter still, it is evident that the sum of their lengths will more nearly equal the length of AXB.

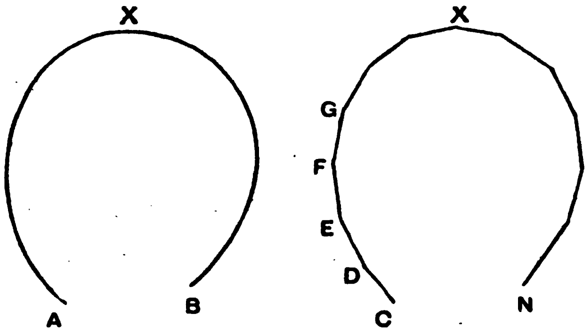

FIG. 6

Expt. 14.—Find the length of the curved line AXB (Fig. 6), **using a pair of dividers** (Fig. 7).

Separate the points P and Q of the instrument by a distance of 4 mm. Place the point P on the end A and turn the dividers until Q also rests on the curve. Now keeping Q fixed, bring round P until it again lies on the curve. The instrument has now stepped along the curve twice, or through a distance of 2 × 4 mm. Continue in this manner until the end B is reached, keeping count of the number of times the dividers have been moved. This number multiplied by 4 mm. gives approximately the length of the curve in mm.

Expt. 15.—Measure the length of the same curve, using a piece of sewing-thread.

Place one end of the thread on A and hold it there with the first finger of the left hand. Stretch the thread along a small portion of the curve, and then hold it with the nail of the first finger of the right hand. Now lift the left-hand finger and place it close to the right-hand one. Removing the latter, again stretch the thread over the next small portion of the line. Continue in this way until the end of the curve is reached. The length of the thread thus used must then be measured by means of a scale.

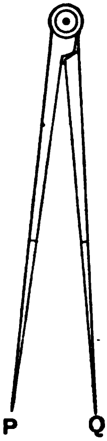

FIG. 7

Expt. 16.—Find the circumference of (1) a penny, (2) a halfpenny, (3) a disc of cardboard, by rolling along a straight line drawn on paper.

Make a fine mark on one face of the disc, as in Fig. 8. Place the mark over the end of the line and roll the disc along it till the mark again touches the line. The circumference is the length of the part of the line passed over. Check your results by rolling along a scale, for example, along the line which joins the ends of all the mm. divisions.

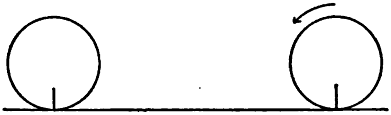

FIG. 8

More accurate determinations of the lengths of curves are obtained by the use of an instrument called the **Opisometer.** It consists of a frame **F** (Fig. 9) attached to a handle **H**, and carrying a pointer **P**. A little wheel **W**, with a milled edge, when turned, moves along a finely cut screw **AB**.

It is used as follows: The wheel is brought to the pointer end of the screw, and then held, by means of the handle, in a vertical position, resting on one end of the curve. The wheel is carefully run along the whole length of the curve and then gently lifted off the paper. **W** will now be some distance from **A**. The wheel is then run *backwards* along a scale until it again comes to a stop at the pointer end of the screw. The distance travelled on the scale is the length of the curved line.

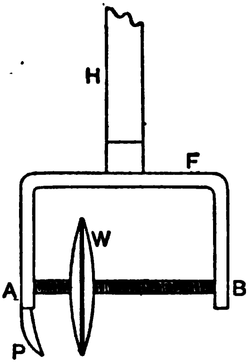

FIG. 9

Expt. 17.—Measure with the opisometer the length of the curve **AB** (Fig. 6), also the length of the curve shown in Fig. 10. In each case make three determinations and take the mean.

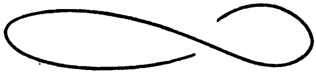

FIG. 10

Expt. 18.—To find how many times the diameter of a circle is contained in the circumference.

Describe three concentric circles of diameters 6, 8, and 10 cm. Measure the length of the circumference of each circle, using (1) the opisometer, (2) a piece of thread. Enter your results in a table drawn up as in the following example:—

Method.	Diameter in cm.	Circumference in cm.	$\frac{\text{Circumference}}{\text{Diameter}}$.
Opisometer	6	18·84	3·140
,,	8	25·16	3·145
,,	10	31·39	3·139

You will notice that the numbers in the last column are very nearly the same, although the circles measured are of different sizes.

If the above experiment be repeated with circles of any convenient diameter, a similar set of numbers will be obtained in the last column.

This points to the fact that **the length of the circumference of any circle divided by that of its diameter is a constant quantity.** That this is the case can be proved by geometry.

This constant ratio is always represented by the Greek letter π (pi) and is equal to 3·14159...

For all practical purposes it is sufficiently accurate if the value of π is taken as 3·14.

Hence **the length of the circumference of a circle** of diameter d or radius r **is πd or $2\pi r$.**

The value of π is often taken as $\frac{22}{7}$. Now $\frac{22}{7} = 3\cdot\dot{1}4285\dot{7}$ which is not quite the same as 3·14159... Hence remember that $\frac{22}{7}$ is only an approximate value of π.

Expt. 19.—Calculate the length of the circumference of a circle whose diameter is 56·23 cm.; also, find the diameter of a circle whose circumference is 33 mm.

Expt. 20.—Measure in three places the diameter of the penny used in Expt. 16. Find the mean value and then calculate the length of the circumference. Repeat this with the halfpenny and the disc of cardboard. Then compare your results with those obtained in that experiment.

Expt. 21.—Closely wrap a tape measure round a cylinder, so that the edges of the tape are parallel to the ends. Note the reading at the point where the tape overlaps. This gives the circumference of the cylinder. Make three or four determinations and find the mean value.

CHAPTER II

MEASURING INSTRUMENTS

Calipers.—The diameter of a cylinder or sphere, the internal diameter of a tube, and many other lengths, cannot be accurately determined by the application of an ordinary scale. In such cases either outside or inside calipers (Fig. 11) may be used.

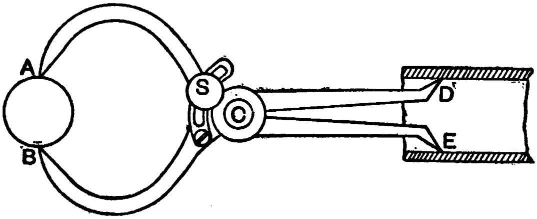

FIG. 11

The figure shows the two kinds combined in one instrument. The outside calipers, ACB, are really dividers with curved legs, and are shown adjusted to measure the diameter of a cylinder. The latter must just pass between the points A and B without forcing them further apart. The legs are fixed in this position by means of a screw, S, and the distance between A and B measured on a scale.

The inside calipers, DCE, are set so that the points D and E just move easily over the internal surface.

Expt. 22.—Measure with the outside calipers the diameter of the cylinder used in Expt. 13.

Expt. 23.—With the inside calipers determine the internal diameter of (1) a beaker, (2) a glass measuring jar, (3) a burette, (4) a piece of large glass tubing.

The vernier.—We have already seen that it is possible to estimate by the eye to a fraction of a scale division; and it will be noticed that if a number of persons make an estimate, it rarely happens that their observations agree.

In order to accurately obtain the fractional part we often make use of a short scale which can slide freely along the graduated edge of the ordinary scale. Such a sliding scale is called a **vernier**,[1] after the name of its inventor. *It enables us to measure to a greater degree of accuracy than is possible with the ordinary scale alone.*

In its simplest form the vernier is divided to measure to tenths of a scale division. A distance equal to nine ordinary scale divisions is divided into ten equal parts to form the vernier scale (Fig. 12).

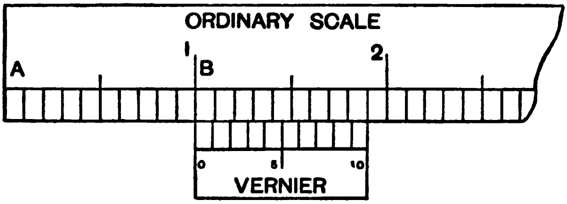

FIG. 12

Each division of the vernier is therefore shorter than a scale division by one-tenth of a scale division. In the figure the zero of the vernier points to the reading 1 on the scale, indicating that the distance **AB** is exactly one inch.

In Fig. 13, however, **AC**, the distance to be measured, is seen to be greater than 1·3 inches and less than 1·4 inches.

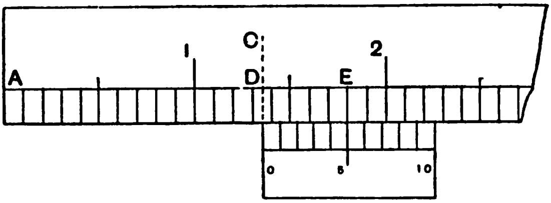

FIG. 13

One person might estimate the distance to be 1·35, and another would possibly say 1·36. The vernier enables us to determine the correct figure to be put in the second decimal place.

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[1] Invented 1631 by Pierre Vernier.

Looking along the vernier we find that the fifth division coincides with a division on the scale. Of course no other vernier division is at the same time coincident with a scale division. Also the five vernier divisions, 0 to 5, or C to E, are shorter than the five scale divisions, D to E, by five-tenths of a scale division or five-hundredths (·05) of an inch. Thus the correct reading is 1·35 inches.

Hence the rule: (1) Read up to the zero of the vernier correct to one of the divisions of the scale; (2) look along the vernier and notice the number of its division which is coincident with a scale division. This number gives the decimal part of the scale division to be added to the reading obtained in (1).

The vernier scale is not always constructed as above, for in some cases twenty divisions correspond to nineteen of the scale. In this instance a reading can be obtained to a twentieth of a scale division, *i.e.* to a two-hundredth of an inch or cm., according to the marking on the scale.

For the measurement of angles the scale is part of a divided circle, and so is the vernier which moves along its edge, but the principle is the same.

Expt. 24.—Construct a scale and vernier in your note-book, as follows:—

Draw a line AB (Fig. 14) 5 inches long. Mark off along it inches and tenths of an inch in the manner shown. This is the scale.

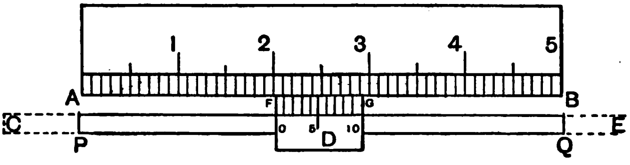

FIG. 14

Draw the outline of the vernier portion, CDE, on a separate piece of paper, making FG exactly nine-tenths of an inch in length. Divide FG into ten equal parts, and draw short lines from the divisions to form the vernier scale. Rub out all construction lines and then cut out the portion CDE.

Lay the vernier with its projecting slips on the edge of the scale, as in the figure, and make two slits, at P and Q exactly below A and B respectively, in the paper on which the ordinary scale is drawn. Pass a projecting piece through each slit. The vernier can then be moved in either direction along the scale. Each slip should be not less than six inches in length.

In Fig. 14 the zero of the vernier points to a reading greater than 2 and less than 2·1.

∴ (1) Reading correct to a scale division = 2·0.

(2) Number of vernier division which coincides with a scale division = 3.

∴ Reading of vernier is 2·03.

Expt. 25.—Place the vernier in different positions along the scale and note the readings. Ask your teacher if each reading is correct.

Also set the vernier to read 1·63, 3·04, 3·4, and 2·97.

The vernier slide calipers.—Fig. 15 illustrates one of its common forms. It consists of a rectangular bar of metal, **S**, upon which a scale is engraved. **A** and **B** are two metal jaws, one of which, **A**, is attached to the bar **S**, and the other, **B**, to the frame **F** which slides along the bar. The sliding piece, which may be clamped in any position by means of the screw **C**, has the vernier scale engraved on it.

When the faces of **A** and **B** are brought into contact the zero of the vernier coincides with the zero on the scale, indicating that the distance between the jaws is nil.

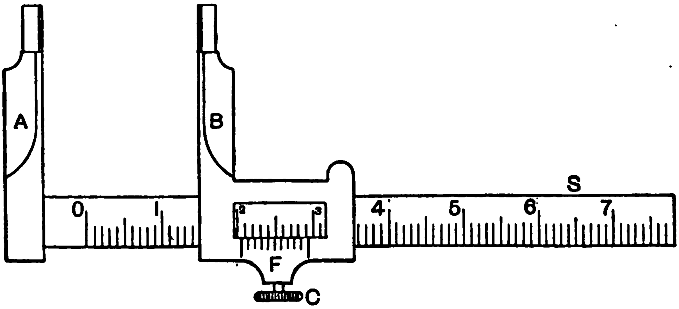

FIG. 15

The object to be measured, a sphere or cylinder for example, is placed between the jaws. **B** is then moved up until in contact with the object. The distance between the jaws, *i.e.* the diameter of the sphere or cylinder, is then found by reading the scale and vernier.

Expt. 26.—Using the vernier calipers, determine the diameter of the sphere, and the diameter and length of the cylinder, employed in previous experiments. In each case make a number of determinations and find the mean value.

Expt. 27.—Measure the diameter of (1) a penny, (2) a halfpenny.

The micrometer screw gauge (Fig. 16).—This instrument consists of a steel frame, **F**, through one end of which passes a small accurately made screw, **S**, of known pitch, usually 1 mm. The screw can be rotated by turning the collar or sleeve, **C**, to which one end of it is fixed. The rotation causes the collar to advance or recede over the cylindrical piece **P** which is fixed to the frame. **A** is a steel plug with planed end, and is fixed in its place by a screw capable of adjustment. The face of the screw **S** is also plane and parallel to that of **A**. There is a linear scale of mm. engraved on **P**, and the bevelled edge of the collar **C** is divided into one hundred equal parts.

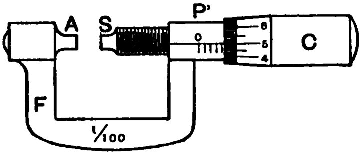

FIG. 16

When the collar makes *exactly one turn*, **S** advances or recedes a distance equal to the pitch or step of the screw, that is, through 1 mm.; and the bevelled edge of **C** moves along the linear scale through a like distance.

When **A** and **S** are in contact the zero mark of the linear scale should coincide with the bevelled edge, and the zero mark on the edge should be a continuation of the line on which the linear scale is marked. Should this be not the case, adjustment must be made by means of the screw which moves the plug **A** in or out as required.

The screw gauge is used to measure the diameter of wires, the thickness of thin sheets of metal, etc.

Expt. 28.—To find the diameter of a piece of wire by means of the micrometer screw gauge.

After adjusting the zero, separate the planed faces, place the wire between them, and turn the collar until the wire is just held by the faces. The number of whole turns, *i.e.* the number of mm., is read off on the linear scale. The hundredths of a turn are obtained by noting the number of the division on the circular edge which is in the same straight line as the line on **P**.

Example. Reading on linear scale, 2 divisions = 2 mm.
Reading on bevelled edge, 43 divisions = ·43 mm.
∴ Diameter of wire = 2·43 mm. = 0·243 cm.

Measure also (1) the diameter of a piece of knitting needle, (2) the external diameter of a glass tube.

Expt. 29.—The diameter of a wire may also be found as follows:—

Carefully wind a good length of the wire, in one layer, round a piece of glass tubing or pencil, so that *each turn is in close contact with its neighbour.* Count the number of these turns and measure the length occupied by them. Divide this distance by the number of turns and so obtain the diameter of the wire.

CHAPTER III

MEASUREMENT OF AREA

In the measurement of the extent of the surface of a body, or its area, the unit used is the area of a square each side of which is of unit length. The metric standard unit of area is thus the square metre. However, just as the centimetre is very often a more convenient unit of length than the metre, so the sq. centimetre is usually employed instead of the sq. metre. Similarly, in the British system we use as units of area a sq. inch, a sq. foot, etc.

Expt. 30.—To find the area of a rectangle.

Draw a rectangle of length 5 cm. and breadth 3 cm. Divide two adjacent sides into centimetres, and through the points of division draw lines parallel to the sides of the rectangle. Then the figure will be divided up into a number of small squares, and the side of each little square will be 1 cm. in length. Thus we have divided the rectangle into sq. centimetres.

Now, as there are three horizontal rows each containing five squares, the number of squares in the figure must be 3 × 5 = 15. Therefore the area is 15 sq. cm.

The area of a rectangle is thus obtained by multiplying the length by the breadth. If these measurements are made in inches, the area will be obtained in sq. inches.

Expt. 31.—Determine the area, in sq. cm., of (1) a page of your note-book, (2) the top of your work-table, (3) a postcard.

Expt. 32.—To find the area of a parallelogram.

Draw on a sheet of paper a parallelogram, **ABCD** (Fig. 17). From **B** draw **BF** perpendicular to **DC**, and from **A** draw a perpendicular to meet **CD** produced in **E**.

Cut off the triangle **BFC** and superpose it on the triangle **AED**. Then, since these triangles are equal, we thus convert the parallelogram **ABCD** into the rectangle **ABFE** without altering in any way the area. Thus the area of the parallelogram **ABCD** is equal to the area of the rectangle **ABFE**.

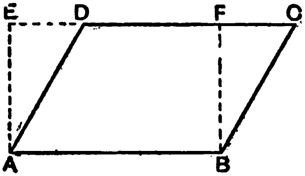

FIG. 17

Hence the area of a parallelogram is obtained by multiplying the base by the perpendicular distance between the base and the opposite side.

Make **AB** = 10 cm. and **BC** = 8 cm., and find the area of the parallelogram.

Expt. 33.—To find the area of a triangle.

Draw on a sheet of paper a triangle **ABC** (Fig. 18). Draw **CD** parallel to **AB**, and **BD** to **AC**; then the figure **ABDC** is a parallelogram whose area is twice that of the triangle **ABC**, because the two triangles **ABC** and **BCD** are equal.

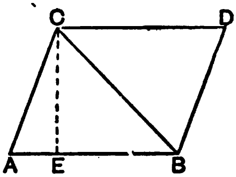

FIG. 18

But the area of the parallelogram is equal to **AB** × **CE**. Hence the area of the triangle **ABC** is equal to half the product of **AB** and **CE**, that is, half the base multiplied by the perpendicular height.

Cut out the triangle **BCD**, and show by actual superposition that it is equal to the triangle **ABC**.

Take in turn each side of the triangle **ABC** as the base, measure the base and the perpendicular height in each case and calculate the area of the triangle.

The area of a regular hexagon.—Divide the hexagon into six equal equilateral triangles as in Fig. 19.

Then area of hexagon

$= 6$ times area of one of the triangles

$= 6 \times (\frac{1}{2}\ AB \times EF)$

$= 6 \times (\frac{1}{2}\ AF \times EF)$

$= 6 \times (\frac{1}{2}\ AF \times \frac{1}{2}\ ED)$

$= \frac{3}{2} \times (AF \times ED)$

$= \frac{3}{4}\ AC \times ED.$

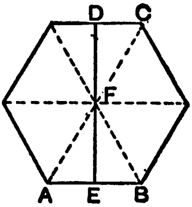

FIG. 19

AC and **ED** are the distances most readily measured when using the slide calipers.

The area of any irregular figure bounded by straight lines may be found by dividing it up into triangles, and then adding together the areas of the triangles.

Area of a circle.—Suppose the circumference of the circle (Fig. 20) divided into a large number of very small parts of lengths s_1, s_2, s_3, etc.; and let the ends of each part be joined to the centre O. Since the little portions are exceedingly short they are practically straight lines, so that the figure is divided into a large number of small triangles.

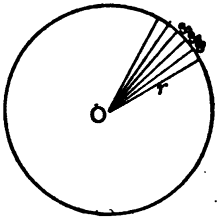

FIG. 20

The area of the circle will therefore be equal to the sum of the areas of these triangles.

The area of the first little triangle

$$= \tfrac{1}{2}\,(\text{base} \times \text{perpendicular height})$$
$$= \tfrac{1}{2}\,(s_1 \times r)$$

since the altitude of each triangle is practically the radius r of the circle.

Hence the area of the circle

$$= \tfrac{1}{2}s_1r + \tfrac{1}{2}s_2r + \tfrac{1}{2}s_3r + \ldots$$
$$= \tfrac{1}{2}r(s_1 + s_2 + s_3 + \ldots)$$

But $s_1 + s_2 + s_3 + \ldots$ etc., is the sum of all the small parts into which the circumference is divided, and is clearly equal to the length of the circumference, that is, $\pi \times$ diameter or $\pi \times 2r$.

$$\therefore \textbf{Area of circle} = \tfrac{1}{2}\mathbf{r} \times 2\pi\mathbf{r}$$
$$= \pi\mathbf{r}^2.$$

If d be the diameter, then $r = \frac{d}{2}$, and the area in terms of the diameter is $\pi\left(\frac{d}{2}\right)^2$ or $\pi\frac{d^2}{4}$.

Expt. 34.—Find the area of a circle whose circumference is 40 cm., also **the diameter of a circle whose area is 95·07 sq. cm.**

Area of an ellipse.—Construct an ellipse as follows: On a sheet of paper fixed on a drawing-board mark two points, A and B (Fig. 21), about 8 cm. apart. Pass a loop of thread over a pencil at P and over two pins fixed into the board at **A** and **B**. Move the pencil over the paper, keeping the thread tight, and thus trace out the ellipse.

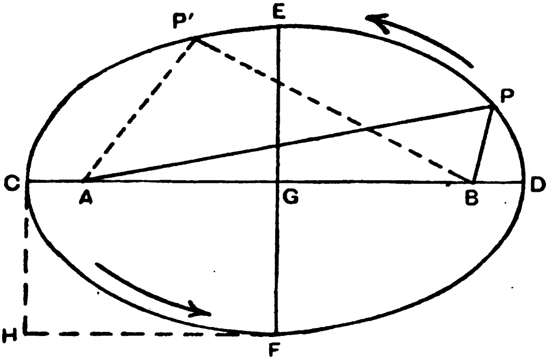

FIG. 21

Produce **AB** to **C** and **D**; bisect **AB** in **G**, and draw **EGF** at right angles to **CD**. Then **CD** is called the *major axis* and **EF** the *minor axis* of the ellipse. The points **A** and **B** are called *the foci*.

The area of the ellipse $= \pi \times \mathrm{EG} \times \mathrm{CG}$

$= \pi \times$ **(product of the semi-axes).**

Thus the area is equal to π times the rectangle **CHFG**, or π times the rectangle contained by the semi-axes.

Expt. 35.—Measure the length of the semi-axes of the above ellipse, and calculate its area in sq. cm.

Expt. 36.—Area of the curved surface of a cylinder.

Cut out a rectangular piece of paper, which, when wrapped round the cylinder, will just cover it. The length of the rectangle is clearly the circumference of the cylinder, whilst the width is equal to the length of the cylinder.

Thus the area of the surface of a cylinder, whose radius is r cm. and length l cm., is $2\pi rl$ sq. cm.

Measure the length and breadth of the rectangle and find its area; also measure the diameter and length of the cylinder by means of the slide calipers, and calculate the area. Compare the two results.

The area of the surface of a sphere of radius r **is** $4\pi r^2$, or 4 times the area of a circle of the same radius.

Measurement of the areas of figures by means of paper ruled into small equal squares.—Suppose any given figure drawn on such paper; then, if we count the number of squares within the figure, and multiply this number by the area of one of them, the product is the area of the figure.

When the bounding line cuts through a number of small squares, the portions of a square within the line must be added together, an estimate being made of the number of whole ones to which they are equal.

The area of one small square—if not known—is obtained as follows: Mark out a large square on the paper, the length of each side being 10 small squares. This large square contains 100 small ones. Measure the large square and calculate its area; then divide the result by 100, to obtain the area of one small square.

Expt. 37.—Determine by the above method the area of a circle 10 cm. in diameter.

Expt. 38.—Using the method of counting squares, prove the following statement: *The areas of triangles on equal bases and between the same parallels are equal.*

Expt. 39.—In a right-angled triangle the area of the square on the hypotenuse is equal to the sum of the areas of the squares on the other two sides.

Test this statement by drawing the figures on squared paper, and then counting the small squares.

Expt. 40.—Draw an ellipse on squared paper, and by counting squares find its area. Also measure its semi-axes, and hence calculate the area.

Expt. 41.—Calculate the area of a circle 8 cm. in diameter, (1) in sq. cm., (2) in sq. ins. Hence find the number of sq. cm. in 1 sq. in.

Expt. 42.—Calculate the area of cross-section of each of the wires whose diameters you measured.

The determination of areas by methods requiring the use of a balance is dealt with in Chapter V.

EXAMPLES II

1. How many sq. mm. are there in 1 sq. metre?
2. How many sq. m. are there in 1 sq. Km.?
3. The diagonal of a square is 250 cm. Find the length of a side and the area of the square.
4. The area of a triangle is 100 sq. mm., and its vertical height is 1·5 dm. Find its base.

5. Find the area of a regular hexagon of side 6 cm.

6. Compare the area of the surface of a sphere of radius r with that of the curved surface of a cylinder of the same radius and length equal to $2r$.

7. The area of a circle is 1 sq. metre. Find its circumference and its radius.

8. Find the radius of a circle of circumference 40 cm.

CHAPTER IV

MEASUREMENT OF VOLUME

The volume of a body is the amount of space it occupies.

For scientific purposes the unit of volume is the volume of a cube each edge of which is 1 cm. It is called a **cubic centimetre**, and is denoted by **1 c.c.**

A **cubic decimetre**, *i.e.* 1,000 c.c., is called a **litre**. This is the unit of volume used on the Continent for measuring wine, milk and liquids generally.

As units of volume in the British system we have the cubic inch, cubic foot, cubic yard, the pint, gallon, etc.; *but the pint and the gallon have no simple relation to the units of length.*

Volume of a right prism.—*A right prism is a body whose ends are parallel and whose sides are everywhere at right angles to the ends.*

The volume of such a body can evidently be found in the following manner: Find the area of the base (Fig. 22) to determine the number of unit cubes which would just cover it and form a layer 1 cm. high. Then measure the height of the prism to find how many such layers would be required to build up the solid.

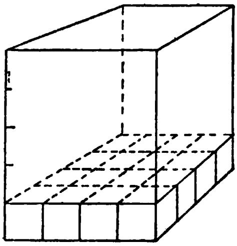

FIG. 22

The total number of unit cubes, *i.e.* the volume of the prism, is then obtained by multiplying the area of the base by the height.

Thus **the volume of a cylinder** of radius r cm. and length l cm., **is $\pi r^2 l$** cubic centimetres.

The above method of finding the volume of a right prism will also apply to the case of a sloping prism such as is represented by B (Fig. 23), the volume being equal to the area of the base multiplied by the vertical height. Hence if A and B represent prisms upon equal bases and between the same parallel planes, their volumes are equal.

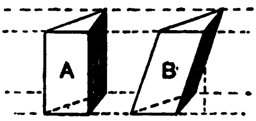

FIG. 23

Expt. 43.—Determine the volume of various right prisms, made of wood or metal, on rectangular, triangular, circular and hexagonal bases. Make the necessary measurements with the slide calipers (1) in inches, (2) in cm. Hence determine the number of c.c. in 1 c. in. Enter your results in a table, as below:—

Name of prism.	Vol. in c.c.	Vol. in c. ins.	Number of c.c. in 1 c. in.
Square Triangular Hexagonal Cylinder			

Volume of a pyramid.—*A pyramid is a body standing on a square, triangular or polygonal base, and terminating in a point at the top.*

We may consider a cube to be made up of six equal pyramids, the apex of each being at the centre point of the cube, and each face of the latter forming the base of a pyramid. The volume of such a square pyramid is equal to one-sixth the volume of the cube. From this it is seen that **the volume is equal to the area of the base multiplied by one-third the vertical height.**

This rule applies to all pyramids.

The volume of a cone may be deduced from that of a pyramid, for a cone may be regarded as a pyramid with an infinite number of faces. Hence the **volume of a cone = $\frac{1}{3}$ area of base × vertical height**

$= \frac{1}{3}\pi r^2 h$.

Expt. 44.—Determine the area of the base and the vertical height of each of the given pyramids. Hence calculate the volume.

Volume of a sphere.—Suppose the surface divided up into a large number of triangular areas, and the centre of the sphere joined to the angular points of each little triangle. In this way we may suppose the sphere split up into an infinite number of pyramids, the height of each being the radius, r, of the sphere.

Then the volume of each little pyramid = area of base $\times \frac{1}{3}r$, and the sum of the areas of the bases is the area of the surface of the sphere.

$$\therefore \textbf{Volume of sphere} = 4\pi r^2 \times \tfrac{1}{3}r = \tfrac{4}{3}\pi r^3.$$

Expt. 45.—Determine the volume of the given sphere.

The diameter should be measured, by means of the vernier calipers, in three or four places, and the mean value used for the calculation.

Expt. 46.—Determination of the volume of a solid by the displacement of water.

Place some water in a cylinder or measuring jar which is graduated in c.c.; note the reading of the surface of the water, and then carefully put in the solid. It displaces a volume of water equal to its own volume, and thus the level rises. Again take the reading at the surface of the water. Obviously the difference of the two readings is the volume of the solid.

Find in this way the volume of the sphere used in the last experiment.

Expt. 47.—Determine in the above manner the volume of an irregular piece of marble, glass rod, iron, sealing-wax, lead or coal. Care must be taken to remove any air-bubbles found sticking to the solid.

Expt. 48.—Measurement of the volume of a solid by means of a burette.

Fix vertically on the outside of a beaker a narrow strip of gummed paper. Make a fine pencil mark on the paper at a distance from the base of the beaker a little greater than the height of the solid. Note the height of the water in the burette, and then fill the beaker up to the mark by running the water from the burette. Again note the reading on the burette. The difference of the two readings is the volume of the beaker up to the mark. Next pour out the water and place the solid in the beaker; then fill up to the mark again with water from the burette, noting the volume required, which will be less than in the first case. The difference between these two volumes gives the volume of the solid.

Determine by this method the volume of a regular solid and of an irregular piece of marble.

The determination of volumes by methods which require the use of a balance is dealt with in the next chapter.

EXAMPLES III

1. How many c.c. are there in 1 cubic metre?

2. How many c.c. are there in 1 c. inch, if there are 2·5 cm. in 1 inch?

3. How many c.c. are there in 143·6 litres?

4. The edge of a cube is 10 cm. Find the area of its surface, its volume, and the length of a diagonal of the cube.

5. Find the volume of a pyramid of 21·9 metres altitude, the base being a square of 10·5 metres side.

6. Find in litres the contents of a box, 60 cm. long, 450 mm. wide and 2 decimetres high.

7. The diameter of the base of a cone is 3·5 cm. and the vertical height is 15 cm. Find the volume.

8. Compare the volume of a sphere of radius r with that of a cylinder of the same radius and length equal to $2r$.

CHAPTER V

MEASUREMENT OF MASS—THE SPRING BALANCE—THE COMMON BALANCE—DETERMINATION OF AREAS AND VOLUMES BY WEIGHING

The mass of a body is the quantity of matter in it. In measuring the mass of a body we therefore compare the quantity of matter in it with that in some body taken as the standard.

Mass must not be confused with volume, for 1 c.c. of iron has a greater mass than 1 c.c. of water.

The British unit of mass is the quantity of matter in a certain piece of platinum kept at Westminster, and is called the **pound avoirdupois.**

The unit of mass in the **Metric** system is called the **kilogram,** and is the mass of a certain lump of platinum made by Borda in 1795. It is kept at Paris.

The unit adopted for scientific use is the **gram.**

For all practical purposes **it is defined as the mass of 1 c.c. of pure water at a temperature of 4° C.**

The divisions of a gram are decimal.

A milligram, or $\frac{1}{1000}$th of a gram = ·001 gm.
A centigram, or $\frac{1}{100}$th of a gram = ·01 gm.
A decigram, or $\frac{1}{10}$th of a gram = ·1 gm.

Gravitation.—Whatever changes, or tends to change, the state of rest or uniform motion of a body, is called *force*. The most commonly occurring force is that of *gravitation*, which is a force of attraction. Sir Isaac Newton, in 1665, made experiments and observations which led him to the conclusion that every body in the universe attracts every other body, and that the force of attraction is proportional to the mass of the body, a large mass exerting a greater attractive force than a small mass. Also, the force is less the greater the distance between the bodies, according to a definite rule.

We have to consider at present only the force of attraction which is exerted between the earth and any body on or near its surface. This force is called the **weight** of the body.

If we hold up a piece of lead on the hand, we notice that it presses downwards, towards the earth's centre, and we express this by saying that it is heavy, it has weight. If the hand is removed, the lead, attracted by the earth, moves towards the ground; and, if it were possible, it would continue to fall until it reached the centre of the earth, where ultimately it would remain at rest and have no weight at all.

The greatest force, or pull, exerted on a body occurs at the earth's surface; and the force diminishes if the body is moved from the surface either towards or from the centre of the earth.

Distinction between mass and weight.—This is of great importance. The mass of a body, *i.e.* the quantity of matter it contains, is invariable wherever the body may happen to be. A pound has the same mass whether at any place on the earth's surface or anywhere else in the universe. The weight of a body, however, is the force with which it is attracted by the earth; and since the earth is not quite spherical, but is flattened at the poles and bulges out at the equator, *the weight of a body is greatest at the poles and decreases as we move towards the equator.*

Also, at the top of a mountain a body weighs less than at sea-level.

The weight of a given quantity of matter is therefore not the same in different latitudes, or at different altitudes.

In order to demonstrate this difference in weight we must measure the weight by means of a delicate spring balance (Fig. 24).

The pull of the earth on a mass of 1 lb. is called the weight of 1 lb. Similarly the force with which the earth attracts a mass of 1 gram is called the weight of 1 gram.

Expt. 49.—Attach, in turn, various masses to the hook of a spring balance, and observe that the pointer moves down the scale a different distance in each case, thus showing that the weights of the masses, as measured by the pull on the spring, are different.

If, however, the pull on the spring is the same in two cases, we know that those masses are equal, and thus **we can measure mass by weight.**

Further experiments on the spring balance are given in Chapter X.

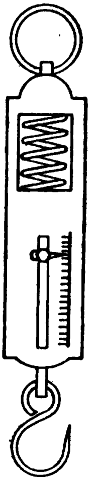

FIG. 24

The common balance.—We have seen that the weight of a body is proportional to its mass, and that two bodies of equal weight are equal in mass. The common balance (Fig. 25) gives us a means of determining when the weights of two bodies are equal.

It consists of a metal beam supported at its centre on a V-shaped axis or knife-edge of steel, about which it is free to oscillate in a vertical plane. A metal pan is suspended from a knife-edge at each end of the beam, and a long pointer is fixed to the centre of

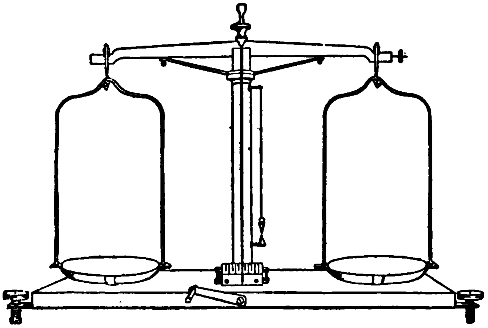

FIG. 25

the beam and moves over a scale as it vibrates. When the balance is in use the knife-edges rest and turn on plates of polished agate. To prevent injury to the knife-edges when the balance is not in use, or when weights are being changed, they are lifted off the agate planes by turning a lever, and the beam then rests on two supporting screws and the pans on the stand.

The box of weights—to be quite correct, we should say the box of masses—employed with such a balance usually contains the following, their values being in grams :—

100, 50, 20, 20, 10, 5, 2, 2, 1,
·5, ·2, ·2, ·1, ·05, ·02, ·02, ·01.

The larger weights are made of brass; the decigrams and centigrams of platinum, aluminium or nickel.

When using a balance the following rules should be observed :—

I. See that the pans are quite clean and free from dust.

II. Gently turn the handle over from the left to the right, and thus raise the beam. See that the pointer, when at rest, is at the centre or zero mark of the ivory scale, or vibrates to the same distance on each side of it. Should this not be the case, adjustment can be made by moving, in the required direction, the small adjusting mass shown at the right-hand end of the beam.

III. Never add or remove a weight, or touch the balance, while the beam is raised, and never lower the beam with a jerk.

IV. Place the body to be weighed in the left-hand pan, and the weights in the right-hand pan. When using a chemical balance always place the substance to be weighed in a counterpoised vessel.

V. Never touch the weights with your fingers; always use the forceps supplied for this purpose.

VI. Place the largest weight which is not too great in the right-hand pan, and then, without missing any, add each smaller weight in turn, leaving it if too light, removing it if too heavy.

VII. When the pointer swings equally on each side of the zero, reckon up the mass by seeing what weights are missing from the box. Check your result on replacing the weights in the box.

VIII. Never place anything on the pan which is likely to injure it, such as anything hot or wet.

IX. Always state the result of a weighing in grams and decimals.

Expt. 50.—Find the mass of a penny, a halfpenny, a piece of chalk, etc.

Expt. 51.—Very carefully determine the mass of an ounce weight. Then fill in the following table:—

1 ounce = . . . gm., 1 pound = . . . gm.
1 gram = . . . oz., 1 kilogram = . . . lbs.

It has been stated that the beam is supported at its middle point, and that the pans hang from the ends of the beam. The distances from the points of suspension of the pans to the centre knife-edge are therefore equal; that is, *the arms* of the balance are equal. The above method of weighing gives the true mass of a body only when the balance is in accurate adjustment, or when the arms are of equal length and the pans of equal mass. Should the arms be of unequal length, the following method of weighing may be employed:—

Method of obtaining the true weight of a body with an ill-adjusted balance by the method of substitution.—A body, called the counterpoise, *heavier* than the object to be weighed, is placed in the left-hand pan, and weights are placed in the other pan until the balance is in equilibrium. The weights are noted and then removed. Suppose they equal W gm. They are replaced by the object whose weight is required and sufficient weights to again balance the counterpoise. Let the additional weights be w gm.

Then, evidently, W gm. = weight of object + w gm., since they both balance the counterpoise under exactly the same conditions. Hence, the weight of the object is $(W-w)$ gm.

The counterpoise may suitably consist of shot, sand or a small piece of lead, brass, marble, iron, etc.

Expt. 52.—Move the small adjusting mass at the end of the right arm of the balance, so as to throw the balance out of adjustment. Then determine, by the above method, the mass of the penny used in Expt. 50.

Another method of obtaining the true weight of a body with a balance having unequal arms is given in the chapter on Simple Machines.

Measurement of area by weighing.

Expt. 53.—To find the area of a given figure, either regular or irregular.

Cut out the figure from a piece of cardboard of uniform thickness, and weigh it. This weight divided by the weight of 1 sq. cm. of the cardboard gives the area of the figure in sq. cm.

To obtain the weight of 1 sq. cm. of the cardboard, cut from it a rectangle, the larger the better, and weigh it. Measure the length and breadth of the rectangle, and calculate its area in sq. cm.

$$\text{Then, weight of 1 sq. cm.} = \frac{\text{Weight of rectangle}}{\text{Area of rectangle}}.$$

Determine by this method the area of (1) a circle of 6 cm. radius, (2) an ellipse, (3) a hexagon of 4 cm. side. Also calculate the areas of the figures, and enter your results in a table, as follows:—

Figure.	Area by weighing.	Area by calculation.
Circle		
Ellipse		
Hexagon		

Expt. 54.—Find the areas of various irregular figures drawn on squared paper which has been pasted on cardboard, (1) by weighing, (2) by the method of counting squares.

Measurement of volume by weighing.

Expt. 55.—To find by displacement of water the volume of a given solid.

Take a test-tube or beaker just wide enough to admit the solid, and fix upright a strip of gummed paper on the outside of the vessel; make a fine pencil mark on the paper at a distance from the base of the vessel a little greater than the height of the solid; then fill with water up to the mark. Now place the vessel on the balance pan—if a test-tube is used, suspend it by the neck with cotton from the hook above the pan—with the solid by its side, and weigh.

Then gently place the solid in the vessel; the water rises above the mark, and the volume of water displaced is, of course, equal to the volume of the solid. By means of a pipette remove the water down to the mark. Now, again weigh the vessel and contents. The result will evidently be less than the weight first obtained by the weight of the water removed. Subtract the two weights, and thus find the weight of water displaced. But every gram of water occupies 1 c.c., therefore the number of c.c. occupied by the solid is the same as the number of grams of water displaced by it. Hence the volume is known.

Note.—To obtain an accurate result the levels must be very carefully adjusted, so that the bottom of the meniscus coincides with the mark.

The solid, which may be either regular or irregular, may conveniently consist of a piece of marble, glass rod or sealing-wax.

In the last experiment you used a vessel which would only just admit the solid. Why? Because it is not a very easy matter to adjust the level of the water *exactly* to the mark; and therefore if you made any little error of adjustment it affects only the smallest possible area. For example, suppose that you ought to have removed another one-sixteenth of an inch of water to bring the level back exactly to the mark. In a test-tube this one-sixteenth of an inch layer of water would weigh very little and the error would be fairly small. But had the solid been placed in a wide-mouthed beaker, this layer would spread over a much larger area and the error would be much greater.

Expt. 56.—Find by the above method the average volume of a lead shot.

Expt. 57.—To find by displacement the volume of a piece of paraffin wax, wood, or other substance which floats in water.

Select a weight which will be sufficient to sink the wax. Tie the two together with thread and find the volume of water displaced by them. Then determine the volume of the sinker, and the difference of these volumes is, of course, the volume of the wax. Any air bubbles that cling to the solids must be removed.

The volume of a piece of a substance which floats in water might be found by displacement of some other liquid *in which it would sink.*

The volume of a liquid may be measured by using variously shaped vessels, such as a **burette**, a **graduated glass cylinder**, a **measuring flask** or a **pipette.** Flasks and graduated glass cylinders measure the volume of the liquid placed in them; whilst burettes and pipettes are used to deliver any required volume of liquid.

Expt. 58.—To find the volume of a bottle or flask.

Make with a file a fine mark on the neck of the bottle, and then weigh the bottle. Fill with water up to the mark, adjusting the level by means of a pipette, and again weigh.

Example. Weight of bottle = 26·65 gm.
Weight of bottle + water = 85·97 gm.
∴ Weight of water = 59·32 gm.
Hence volume of bottle up to mark = 59·32 c.c.

CHAPTER VI

STATES OF MATTER

Matter is that which affects our senses. It can exist in three states: (1) The solid state, (2) the liquid state, (3) the gaseous state, and it has very different properties in each state.

Liquids and gases are classed together as **fluids.**

Many forms of matter may exist in all three states under ordinary conditions. Thus the same form of matter can exist in the solid state as ice, in the liquid state as water, and in the gaseous state as steam.

Matter **occupies space,** has **weight** and possesses **inertia,** that is, it requires the application of a force to set it in motion. It is also **impenetrable, porous** and **indestructible.** By saying that it is impenetrable we mean that two portions of matter cannot occupy the same space at the same time.

Porosity.—By saying that matter is porous we mean that bodies do not fill the whole of the space which they appear to occupy. There are empty spaces within them. The particles of which a body is composed are separated by minute spaces or pores. Some substances, such as sponge, cork and pumice, have visible holes in them. In other bodies, such as lead, iron, gold, etc., the interstices cannot be seen with the naked eye, but their existence may be proved by forcing water through them. Thus Francis Bacon forced water through lead by squeezing and hammering a leaden sphere full of water.

Hydrogen will pass through heated platinum, and carbonic oxide through hot iron. Blotting paper and filter paper are obviously porous, while the porosity of liquids is usually shown by the following experiment:—

Expt. 59.—About half fill a burette with water, and then gently add sufficient alcohol to bring the level up to the top graduation mark. Shake the liquids so as to thoroughly mix them, and then note the volume of the mixture. It will be found that a diminution in volume has taken place, the reason being that the liquids have filled up interstices between the particles which were previously empty.

When a piece of chalk is thrown into water, bubbles of air are seen to rise to the surface, being expelled by the water from the pores in the chalk.

A common method of cleaning mercury is to filter it by squeezing it through a piece of chamois leather. Also, a beaker quite full of warm water may be made to hold still more matter, for with care a large quantity of sugar may be dissolved in the water without causing any overflow.

Matter also possesses the properties of **cohesion, compressibility** and **elasticity**.

Cohesion is the attractive force which holds together the particles of a body.

Compressibility is the property in virtue of which the volume of a body can be made smaller by pressure. It follows as a natural consequence of porosity, for by applying pressure the particles can be made to go closer together. Gases are very compressible, whilst liquids are almost incompressible. Usually, in the case of solids very great pressure has to be exerted to cause any diminution in volume.

Elasticity.—An elastic body is one which requires the application of force to change its shape or volume, and which springs back and regains its original shape and volume as soon as the force is removed.

Gases and liquids have only elasticity of volume, there is no elasticity of form or shape; they recover their original volume after compression, when the compressive force is removed.

Solids possess volume-elasticity—*i.e.* they resist changes of volume; but they also have elasticity of form. Resistance to change of shape or form, the volume being supposed to remain constant, is called **rigidity**.

We recognise the three states of matter by the following characteristics:—

I. The solid state.

A solid has a *definite shape and size*, to any change of which it offers considerable resistance, that is, cohesion is strong in a solid.

We have seen that solids possess elasticity of volume; thus, a sphere of glass or ivory when let fall is slightly compressed on touching the ground; recovery immediately takes place, causing the sphere to rebound. Thus, when we say that glass and ivory are very elastic, but that putty, wet clay and dough are not, we have in view elasticity of volume.

Expt. 60.—Slightly grease a smooth metal plate, or a slab of marble, and allow two balls, one of glass and the other of ivory, to fall upon it. Notice the well-marked circles made on the grease, proving that the spheres have been compressed.

But solids offer resistance to any alteration of their shape by *stretching*, *bending* or *twisting;* they possess elasticity of form.

If a long, fine metal wire be stretched by fixing it at one end and hanging weights to the other, it will slowly go back to its original length when the weights are removed, provided that it has not been stretched too far; for many substances are only elastic within certain limits, called *the limits of elasticity.*

If one end of a steel strip or a piece of whalebone is fixed in a vice and the other end is drawn to one side and then let go, the strip vibrates backwards and forwards, finally coming to rest in its original position. A strip of lead treated similarly remains bent after the removal of the bending force, the particles making no effort to recover their former positions.

The vibrating prongs of a tuning-fork also illustrate elasticity brought into play by bending.

Again, when a cylindrical rod or wire fixed at one end and weighted at the other is twisted round through a certain angle and then let go, the wire untwists back to and beyond its starting point, and then returns, in this way twisting and untwisting until it comes to rest.

Glass, indiarubber, ivory, whalebone and steel are solids which possess elasticity in the highest degree.

Solids possess **hardness**, **tenacity**, **ductility** and **malleability**.

Hardness means resistance of the particles of a solid to any change in their relative positions, such as would be produced were the body scratched. Diamond is the hardest known substance, for it will scratch all bodies and is not scratched by any. The following table gives a few substances in the order of their hardness:—

1. Diamond.
2. Sapphire.
3. Topaz.
4. Quartz (rock-crystal or flint).
5. Glass.
6. Steel.
7. Iron.
8. Copper.
9. Lead.
10. Wax.

The **tenacity** of a substance means the resistance of its particles to complete separation or rupture. It is usually measured by hanging weights from a rod or wire of the substance, and observing what force is required to break it. Steel, wrought iron, brass and copper have great tenacity, whilst that of lead is very small.

Ductility is the capacity of a solid for being drawn out into a wire. The most ductile metal is platinum. Gold, silver, copper and iron are also very ductile, whilst zinc and tin are hardly ever seen in the form of wire. Lead is the least ductile of all common metals.

Malleability is the property by virtue of which solids may be hammered out into thin plates. The malleability of gold is remarkable, for a leaf of gold can be beaten out until only $\frac{1}{300,000}$ of an inch in thickness. Such a leaf is actually transparent, transmitting a beautiful green light. Silver, platinum, lead, zinc and copper are also very malleable.

II. The liquid state.

A liquid has *no definite shape*, but always takes the shape of the vessel which contains it. It has, however, under the same conditions, *a definite size or volume*, no matter how much its shape may be altered.

If introduced into an empty vessel of larger volume than itself, it occupies only a portion of the vessel equal to its own volume, and shows a distinct boundary or surface at the top, as water does in a beaker. The surface of the liquid at rest is also *horizontal*. A gas under similar conditions will spread throughout the whole of the vessel, and will show no *free surface*.

Liquids are practically *incompressible*, the application of a very large force producing an extremely small change of *volume*. They yield to the action of the smallest force tending to change their *shape;* that is, they have no rigidity; and if not held by the sides of a vessel they flow, *i.e. the force of cohesion is very small.*

All liquids do not flow with the same facility, for the particles of a liquid experience when they slide over one another a certain amount of resistance, called **viscosity** or fluid friction. Water and alcohol have small viscosity, while oil and treacle are very viscous.

There is no definite line of demarcation between solids and liquids, for some substances occur in an intermediate state, for instance, pitch, honey and tar. They show the general properties of liquids, though very imperfectly.

When a number of communicating vessels contain the same liquid, the level is the same in all (Fig. 26).

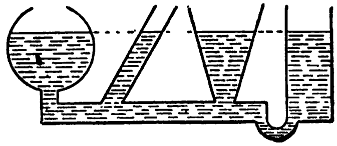

FIG. 26

Fluids transmit pressure equally in all directions.

Pressure is measured by the force acting on unit area; for instance, if a force equal to the weight of 6 lbs. acted on an area of 2 sq. ins., the pressure on the area would be 3 lbs. per sq. in. Let Fig. 27 represent a vessel filled with a fluid, say water, and having openings of equal area at **A**, **B** and **C**, the openings being fitted with water-tight pistons.

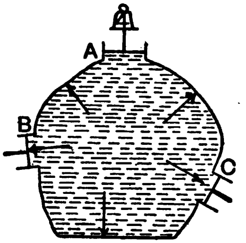

FIG. 27

Let the area of each piston be 1 sq. in., then, if a weight of 4 lbs. be placed on the piston **A**, a pressure of 4 lbs. per sq. in. will be transmitted in all directions, and every sq. in. of the internal surface will experience it. To prevent **B** and **C** moving out, a force of 4 lbs.-weight must be applied to each. If the area of **C** were 2 sq. ins., a force of 8 lbs.-weight would be required to keep the piston in its place.

Liquids can be formed into drops which run together when brought into contact.

III. The gaseous state.

A gas has *neither definite shape nor definite size;* it fills and takes the shape of the containing vessel. *No cohesion at all is shown.*

Gases may be distinguished from liquids by their great *compressibility and expansibility.* They can very easily be compressed into a much smaller space; and when allowed to expand they tend to do so indefinitely, and thus do not show a free surface.

CHAPTER VII

DENSITY

Density.—We often hear the statement made that one substance is *heavier* than another; for instance, lead is heavier than cork. Yet it is quite possible to get a large lump of cork which will weigh more than a small piece of lead. What, then, do we mean when we say that lead is a heavier substance than cork? We mean *that if we were to take a certain volume of lead and an exactly equal volume of cork, the lead would weigh more than the cork;* we mean *that lead is heavier than cork* **bulk for bulk.** This is expressed by saying that lead is denser than cork.

Thus, equal volumes of different substances have different weights and therefore different masses.

From this we infer that the particles of matter in the denser substances must be more closely packed together than in the less dense substances.

The density of any substance is the mass per unit volume of that substance.

If the mass of a substance is M gm., and its volume V c.c., then its density D is $\frac{M}{V}$ gm. per c.c., or $\mathbf{D} = \frac{\mathbf{M}}{\mathbf{V}}$.

Thus the density of water is 62·5 lbs. per cu. ft.
or 1,000 oz. per cu. ft.
or 1 gm. per c.c.

Meaning that 1 cu. ft. of water has a mass of 62·5 lbs. or 1,000 ozs., and that 1 c.c. of water has a mass of 1 gm.

The density of a substance is therefore expressed by different numbers according to the units of mass and volume used. Therefore when giving the density of a substance the units of mass and volume must be stated.

Expt. 61.—To find the density of various regular and irregular solids, such as prisms, spheres or pyramids of wood, and pieces of glass rod, twisted copper wire, brass, iron and zinc.

First weigh the solid to obtain its mass. Next find its volume—by measurement and calculation if a regular solid, by displacement of water if an irregular solid. Then the weight divided by the volume gives the mass per unit volume or the density.

Example. Wt. of a piece of marble = 29·06 gm.
Volume of the marble = 10·70 c.c.

$$\therefore \text{Density of marble} = \frac{29{\cdot}06}{10{\cdot}70}, \text{ or } 2{\cdot}71 \text{ gm. per c.c.}$$

Expt. 62.—To find the density of a liquid by means of a flask.

Make a fine file mark on the neck of the flask. Weigh the flask, fill it with the liquid up to the mark, and again weigh. The difference of the two weights will be the mass of the liquid.

To determine the volume of the liquid, empty the flask and rinse it out with water; then fill it with water up to the mark and weigh it. The resulting weight, minus the weight of the empty flask, will be the mass of the water. From this the volume of the flask up to the mark is known, since 1 gm. of water occupies 1 c.c. Thus the volume of the liquid is obtained. Hence the density can be found.

Example. Wt. of empty flask = 17·47 gm.
Wt. of flask + liquid *A* = 77·72 gm.
∴ Wt. of liquid *A* = 60·25 gm.
Wt. of flask + water = 67·45 gm.
∴ Wt. of water = 49·98 gm.
∴ Volume of flask up to mark = 49·98 c.c.

$$\text{Hence, density of liquid } A = \frac{60{\cdot}25}{49{\cdot}98}, \text{ or } 1{\cdot}20 \text{ gm. per c.c.}$$

Find, by the above method, the density of a common salt solution, and of a saturated solution of copper sulphate, entering your results as shown in the above example.

Pressure due to a column of liquid.—When a small metal disc is held over a hole made in the side or bottom of a vessel filled with a liquid, a pressure is felt against the disc. The liquid then exerts a pressure on the internal surface of the vessel. Also, if a solid is suspended in a liquid, it is subjected to pressure on its outside surface. When the solid is removed its place is taken by an equal volume of the liquid, which experiences *the same pressure of the surrounding liquid.* Thus, at any point in a liquid there is a pressure due to the weight of the liquid. The pressure is proportional to the density of the liquid and to the head of the liquid at the point (the depth of the point below the free surface).

Expt. 63.—Take a hollow glass cylinder, open at both ends, and having one end ground smooth. Close the smooth end by a thin, light, flat disc, which may be held in place by means of a string attached to its centre. Now push the cylinder and disc into water, always keeping them in a vertical position. Let go the string and the disc will be found to remain in position, being kept there by the upward pressure of the water. Pour water very slowly and carefully into the cylinder, and notice that the disc remains fixed until the water inside reaches a level a little below that outside the cylinder.

This experiment shows that the upward pressure on the disc is equal to the downward pressure due to the weight of the column of water inside the cylinder.

Expt. 64.—To compare the densities of two liquids which do not mix, say water and turpentine, by means of a U-tube.

Support the U-tube vertically. Pour the *heavier* liquid (water) in first, and then gently pour the other liquid on to it, taking care not to drive the surface of separation of the liquids into the bend. Suppose the final levels are as in Fig. 28, where the surface of separation is at A. Suppose a horizontal line, AB, drawn through this surface; then the downward pressure at A due to the column AC of turpentine must be the same as the upward pressure at A due to the column, BD, of water, since the surface of separation is at rest. But in each case the pressure is proportional to the head of liquid and to the density of the liquid; therefore

Height AC of turpentine × density of turpentine
= Height BD of water × density of water;

or, $$\frac{\text{Density of turpentine}}{\text{Density of water}} = \frac{\text{Height BD of water}}{\text{Height AC of turpentine}}.$$

Hence the density of turpentine is found, since that of water is unity.

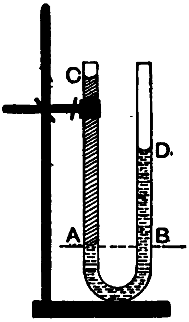

FIG. 28

Thus **the densities of the liquids are inversely as the lengths of their columns above the surface of separation.**

Compare the densities of mercury, glycerine and paraffin oil with that of water.

In the above experiment **the diameters of the two limbs of the tube may be unequal and the method still holds good.**

Also, the arrangement of liquids behaves exactly like an ordinary balance, one scale-pan being supposed to be at A and the other at B. When the masses on the pans are equal, the pans are on the same level. Hence, in the figure, the weight of the column, AC, of turpentine is equal to the weight of the column, BD, of water.

Expt. 65.—To compare by means of the U-tube the densities of two liquids which mix.

First place some mercury in the tube; it will stand at the same level in the two limbs. Mark this level on one limb by means of a piece of gummed paper. Nearly fill one limb with the *less dense* of the two liquids, taking care not to drive the mercury right round the bend; then pour the other liquid into the other limb till the mercury comes back to the mark, *i.e.* to the same level in each limb. Then the weight of the liquid above the mercury in the right limb is equal to the weight of the other liquid above the mercury in the left limb. Therefore

$$\frac{\text{Density of liquid in right limb}}{\text{Density of liquid in left limb}} = \frac{\text{Height of column in left limb}}{\text{Height of column in right limb}}.$$

Compare in this manner the densities of alcohol, milk, saturated solutions of common salt and copper sulphate, with that of water.

Expt. 66.—To compare the densities of two liquids by means of Hare's apparatus.

The apparatus consists of two long glass tubes connected at the top to a T-piece by short pieces of rubber tube, or it may be all in one piece (Fig. 29). The short branch has attached to it a piece of tubing fitted with a clip. The lower ends of the tubes dip into beakers containing the two liquids. By suction at E draw up the liquids into the tubes till one liquid nearly reaches the top, and then close the clip. The pressure inside the tube at B is the same as that at D; also the pressure outside the tube at A is the same as that at C. Therefore

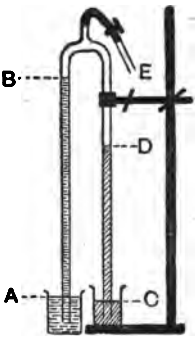

FIG. 29

$$\frac{\text{Density of liquid in AB}}{\text{Density of liquid in CD}} = \frac{\text{Height of column CD}}{\text{Height of column AB}}.$$

Compare by this method the densities of the liquids used in the last experiment with that of water. Enter your results as in the following example:—

Density of salt solution.

Height of salt solution = 28·0 cm.
Height of water = 33·2 cm.

$$\frac{\text{Density of salt solution}}{\text{Density of water}} = \frac{\text{Height of water}}{\text{Height of salt solution}}$$

$$= \frac{33{\cdot}2}{28{\cdot}0} \text{ or } \mathbf{1{\cdot}18}.$$

∴ (since density of water is 1 gm. per c.c.)

Density of salt solution = **1·18 gm. per c.c.**

Note.—The apparatus must be fixed *vertically*, and the two limbs, as also in the case of the U-tube, should be *parallel*. The heights of the columns are measured from the surfaces in the tubes to the surfaces in the beakers. *The levels in the beakers need not necessarily be the same.*

EXAMPLES IV

1. Find the weight of a copper bar whose length is 5 decimetres and sectional area 64 sq. cm. (density of copper is 8·88 gm. per c.c.).

2. A rectangular block of lead whose length is 60 cm., width 500 mm., and depth 2·5 dm., weighs 855 kgm. Find its density.

3. A U-tube contains two liquids that do not mix; the height of the denser above their common surface is 5 inches, and the difference of their levels is 3 inches. Compare their densities.

4. Water and oil (density ·88 gm. per c.c.) are poured into a U-tube, and the water rises 11 inches above the common surface. To what height does the oil rise?

5. A piece of lead weighs 4·56 lbs.; find the weight of a piece of copper of three times the size. Density: lead, 11·4 gm. per c.c.; copper, 8·8 gm. per c.c.

CHAPTER VIII

SPECIFIC GRAVITY—THE PRINCIPLE OF ARCHIMEDES

Specific gravity or relative density.—The specific gravity of a substance is the ratio of the weight of any volume of the substance to the weight of an equal volume of water. Or,

$$\text{Sp. gr. of substance} = \frac{\text{Weight of any volume of the substance}}{\text{Weight of an equal volume of water}}.$$

Suppose we take 1 c.c. of the substance, then

$$\text{Sp. gr. of substance} = \frac{\text{Wt. of 1 c.c. of substance}}{\text{Wt. of 1 c.c. of water}}$$

$$= \frac{\text{Density of substance}}{\text{Density of water}}.$$

Thus, *the specific gravity of a substance is the density of the substance relative to that of water.*

Also, in the Metric system, the density of water is unity (1 gm. per c.c.), therefore **the specific gravity of a substance is represented by the same number as the density when the latter is measured in metric units.** For example, the sp. gr. of mercury is 13·59, and its density is 13·59 gm. per c.c.; the sp. gr. of lead is 11·4, and its density 11·4 gm. per c.c. If, however, the densities were expressed in British units, we should have

Sp. gr. of mercury = 13·59.
Density of mercury = 849·4 lbs. per cu. ft.
Sp. gr. of lead = 11·4.
Density of lead = 712·5 lbs. per cu. ft.

We can now find the weight of a body whose volume is V and specific gravity S.

Let W be the required weight and w the weight of unit volume of water. Then Vw would be the weight of the body if composed of water. But it is made of a substance which is S times as heavy as water; therefore

$$\mathbf{W = VwS}$$

Example 1.—Find the weight of 2 cu. ft. of lead of sp. gr. 11·4.

Here w = the wt. of 1 cu. ft. of water, *i.e.* 62·5 lbs.

Then $W = VwS = 2 \times 62{\cdot}5 \times 11{\cdot}4$, or 1,425 lbs.

Example 2.—Find the weight of 2 litres of a liquid whose sp. gr. is 1·05.

Here w is taken to be the wt. of 1 c.c. of water, *i.e.* 1 gm.

Then $W = VwS = 2{,}000 \times 1 \times 1{\cdot}05$, or 2,100 gm.

The specific gravity of a liquid may be found by the method employed in Expt. 62. In the example there worked out we have

Wt. of liquid A up to the mark = 60·25 gm.
Wt. of water up to the mark = 49·98 gm.

But these are the weights of equal volumes of the two liquids

$$\therefore \text{Sp. gr. of liquid } A = \frac{60{\cdot}25}{49{\cdot}98} = 1{\cdot}20.$$

The equal volumes were obtained by filling up to a certain mark on the neck of the flask. **The specific gravity bottle** (Fig. 30) gives us a means of more accurately measuring equal volumes. It consists of a small glass bottle accurately fitted with a ground glass stopper, which has a fine hole bored through it.

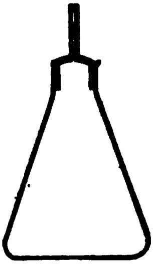

FIG. 30

Expt. 67.—To determine the specific gravity of a liquid by means of the specific gravity bottle.

Weigh the bottle and stopper. Fill the bottle with the liquid and gently insert the stopper, thereby forcing some of the liquid through the hole and leaving the bottle and stopper quite full. Carefully dry the outside of the bottle and weigh. The difference of the two weights is the weight of the liquid. Empty the bottle, rinse it out with water, fill with water, and again weigh. From the result subtract the weight of the bottle and stopper to obtain the weight of contained water. You have thus found the weight of liquid to fill the bottle, and also the weight of the same volume of water; hence the specific gravity of the liquid can be found.

Determine the specific gravities of milk, alcohol, a common salt solution and a copper sulphate solution.

The Principle of Archimedes is so-called because it was discovered by **Archimedes**, a philosopher of Syracuse, in Sicily, where he lived about 250 B.C. The principle is stated thus:—

When a solid is immersed in a fluid it loses a portion of its weight equal to the weight of the fluid which it displaces. Or, in other words, **when a solid is partly or wholly immersed in a fluid, it is pressed upwards with a force equal to the weight of the volume of fluid displaced by it.**

Hence, when a body is weighed in air, it is pressed upwards with a force equal to the weight of the air the body displaces. In consequence, **bodies weigh more in a vacuum than in air.**

The loss of weight in water may be shown as follows:—

Expt. 68.—Suspend a piece of lead, or other heavy substance, from a spring balance, and note the reading on the scale. Then allow the substance to hang completely immersed in water and again note the reading. It will be found to be less than before, showing that the water exerts an upward pressure on the substance.

The truth of Archimedes' Principle may be shown by the following experiment:—

A hollow cylinder, B (Fig. 31), is suspended from the hook which carries the left-hand pan of a balance, and underneath it is suspended a solid cylinder, C, *which exactly fills the inside of the hollow cylinder.* The volume of C is thus equal to the internal volume of B. The cylinders are counterpoised. A jar of water is then placed upon a wooden balance bridge, A, which spans the pan and rests on the base of the balance or on the table, so that the solid cylinder is completely immersed. The water exerts an upthrust on C, which causes it to appear to lose weight. This arm of the balance rises. By means of a pipette B is now filled with water, and equilibrium is restored when it is just full. This proves that the upthrust or loss of weight is equal to the weight of the water required to fill the bucket, B, that is, to the weight of the volume of water displaced by C.

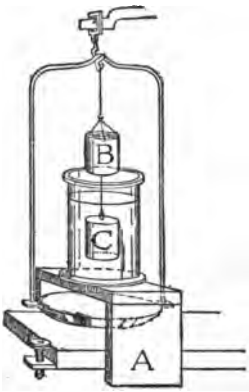

FIG. 31

Repeat the experiment using another liquid in place of water.

Expt. 69. — To prove experimentally Archimedes' Principle.

(1) Weigh a solid, such as a piece of marble or thick glass rod, in the ordinary way, that is, in air. Let its weight be *W* gm. Then suspend the solid, B (Fig. 32), from the hook by means of a single thread of cotton so as to be completely immersed in water. Weigh it. Let its weight be *w* gm. Then, upthrust or loss of weight in water $= (W - w)$ gm.

(2) Find the weight of the water displaced by the solid by the method of Expt. 55. Let the weight be w_1 gm.

It will then be found that $(W - w)$ gm. $=$ w_1 gm., thus proving Archimedes' Principle when water is the fluid used.

Perform the above experiment again, using a solution of copper sulphate instead of water.

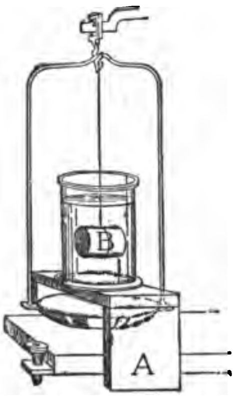

FIG. 32

Expt. 70.—To determine the specific gravity of a solid which sinks in water and is insoluble in it.

Weigh the solid in air and also in water, as in Expt. 69, Fig. 32. The loss of weight in water is equal to the weight of the water displaced.

Let W = the weight in air, w = the weight in water; then $W - w$ = wt. of water displaced, or the wt. of an equal volume of water.

$$\therefore \text{Sp. gr. of solid} = \frac{W}{W - w} \text{ or } \frac{\text{wt. in air}}{\text{wt. in air} - \text{wt. in water}}.$$

$W - w$ is the weight of a volume of water equal to that of the solid. But we know the volume occupied by a given weight of water, and hence the volume of the solid is easily obtained.

Note.—(1) The solid must be wholly immersed throughout the process of weighing. (2) It must not be allowed to touch the sides of the vessel containing the water. (3) A single thread should be used to suspend the solid. (4) Any air bubbles that cling to the solid must be removed.

Determine in the above manner the specific gravities and volumes of a penny, and of pieces of coal, lead, iron, marble, etc.

Expt. 71.—To find the specific gravity of a solid which floats in water.

In this case a heavy weight, such as a piece of lead, called a sinker, must be attached to the solid when the latter is to be weighed in water.

First weigh the solid alone in air; then attach the sinker and weigh in water. Next weigh the sinker alone in water. Let W = wt. of solid in air, w_1 = wt. of sinker alone in water, w_2 = wt. of sinker + solid in water.

Since the solid is lighter than water, it will tend to buoy up the sinker, and therefore w_1 will be greater than w_2.

Let w = wt. of sinker in air, then

$W + w$ = wt. of solid + sinker in air;

$\therefore (W + w) - w_2$ = wt. of water displaced by solid + sinker;

also $w - w_1$ = wt. of water displaced by sinker alone;

$\therefore [(W + w) - w_2] - (w - w_1)$ = wt. of water displaced by solid alone.

$$\therefore \text{Sp. gr. of solid} = \frac{W}{[(W + w) - w_2] - (w - w_1)} = \frac{W}{W - w_2 + w_1}.$$

Hence *only three weighings are required, for it is not necessary to know the weight of the sinker in air.*

Determine in this way the specific gravity of paraffin wax, cork, wood (pine, oak, etc.), and enter your results as follows:—

Weight of substance in air = 5·27 gm.
Weight of sinker in water = 7·67 gm.
Weight of sinker + solid in water = 7·45 gm.

$$\therefore \text{Sp. gr. of substance} = \frac{5{\cdot}27}{5{\cdot}27 - 7{\cdot}45 + 7{\cdot}67} = {\cdot}96.$$

Expt. 72.—To find the specific gravity of a solid soluble in water.

In this case the solid must be weighed in a liquid of known sp. gr., and in which it will not dissolve.

Let W = wt. of solid in air, w_1 = wt. of solid in the liquid, and s = sp. gr. of the liquid.

Then, $W - w_1$ = wt. of the liquid displaced; but since the liquid is s times as heavy as water, volume for volume, $\frac{W - w_1}{s}$ will be the wt. of water which the solid would displace if it were insoluble in it.

$$\therefore \text{ Sp. gr. of solid} = \frac{W}{\frac{W - w_1}{s}} = \frac{Ws}{W - w_1}.$$

Find the sp. gr. of a crystal of copper sulphate by weighing it in a saturated solution of copper sulphate, and of a piece of loaf sugar by weighing it in petroleum.

Expt. 73.—To find the sp. gr. of a powder which is insoluble in water, by means of a sp. gr. bottle.

Weigh the bottle. About one-third fill it with the powder and again weigh. These weighings will give the weight of the powder. Then add water to fill the bottle and shake it in order to remove air bubbles; insert the stopper and weigh. The weight of the water above the substance can now be obtained. Empty the bottle, rinse it out, fill with water, and again weigh.

Let w = wt. of empty bottle,
w_1 = wt. of bottle + powder,
w_2 = wt. of bottle + powder + water,
w_3 = wt. of bottle + water required to fill it.

Then, $w_1 - w$ = wt. of powder,
$w_3 - w$ = wt. of water required to fill the bottle,
$w_2 - w_1$ = wt. of water above the powder.
$\therefore (w_3 - w) - (w_2 - w_1)$ = wt. of water displaced by powder.

$$\therefore \text{ Sp. gr. of powder} = \frac{w_1 - w}{(w_3 - w) - (w_2 - w_1)}.$$

Determine in this way the sp. gr. of fine sand, powdered glass, small brass or iron nails, small lead shot, fragments of stone, etc.

Expt. 74.—To determine the sp. gr. of a liquid by weighing a solid in it, in air, and in water.

Let W = wt. of solid in air, w = wt. in water, and w_1 = wt. in liquid.

Then the wt. of liquid displaced by the solid = $W - w_1$,
and the wt. of water displaced by the solid = $W - w$.

But these are the weights of equal volumes.

$$\therefore \text{ Sp. gr. of liquid} = \frac{W - w_1}{W - w}.$$

Find the specific gravities of common salt solution, milk, turpentine and alcohol, by weighing a piece of thick glass rod or twisted copper wire in them.

Expt. 75.—To find the diameter of a thin wire.

Take a piece of copper or german silver wire 1 metre in length. Wind it into a spiral and then weigh it, first in air and then in water. The weight in gm. of the water displaced gives the volume in c.c. of the wire. The wire is really a very long cylinder, and therefore $\frac{\text{volume}}{\text{length}}$ = area of cross-section of the wire $= \frac{\pi d^2}{4}$ if d is the diameter of the wire. Hence d can be calculated.

Specific gravity of substances by flotation.—We have seen that when a body is partly or wholly immersed in a fluid it is pressed upwards with a force equal to the weight of the fluid displaced. Also, the weight of the body acts vertically downwards. Thus, when a body is placed in a fluid, it comes to rest in such a position that these two opposing forces just balance each other. Therefore, if the weight of the body is greater than the weight of the displaced fluid, the body will sink. It will rise if its weight is less than that of the fluid displaced. If the weight of the body is just equal to the weight of its own volume of the fluid, then it will float anywhere in the fluid. In the following experiments the principle to be remembered is that **the weight of a floating body is equal to the weight of the liquid which it is displacing.**

Expt. 76.—To find by flotation the sp. gr. of a rectangular wooden block.

Very slightly oil the block and place it in a glass vessel containing water. Measure the depth to which it sinks at each corner, and take the mean as the depth of immersion.

Let V c.c. = the volume of the body, V_1 c.c. = the volume of the part immersed, *i.e.* the volume of water displaced, and S = the sp. gr. of the body.

Then, the wt. of the body = the wt. of water displaced;

$$i.e.\ (V \times S)\ \text{gm.} = V_1\ \text{gm.}$$

$$\therefore\ S = \frac{V_1}{V} = \frac{\text{Volume immersed}}{\text{Whole volume}}.$$

In this case, let x = height of block, y = mean depth of immersion, and A = area of upper or lower face of block, then

$$S = \frac{V_1}{V} = \frac{A \times y}{A \times x} = \frac{y}{x}.$$

Therefore the sp. gr. of the block is got by dividing the depth immersed by the whole depth.

Calculate the weight of water displaced, and compare it with the weight of the dry block.

Expt. 77.—To find the sp. gr. of a liquid by flotation.

Close a piece of glass tube, and place in it small shot, fine sand, or a little mercury, so that it will float upright. A paper scale may also be placed inside, in order to measure the depth to which the tube sinks. Then close the other end. A narrow test-tube closed with a cork may be used instead.

Float the tube in water and then in the liquid whose sp. gr. is required. Note the length immersed in each case. Let l_1 cm. = length in water, and l_2 cm. = length in the liquid. Then, if a sq. cm. is the sectional area of the tube, and s the sp. gr. of the liquid, the volume of water displaced $= (l_1 \times a)$ c.c.; its wt. $= (l_1 a)$ gm., and the volume of liquid displaced $= (l_2 \times a)$ c.c.; its wt. $= (l_2 as)$ gm.

But, the weight of the tube = the weight of water it displaces,
and also = the weight of liquid it displaces;

$$\therefore\ l_1 a = l_2 as.$$

$$\therefore\ s = \frac{l_1}{l_2} = \frac{\text{Depth immersed in the water}}{\text{Depth immersed in the liquid}}.$$

Determine in this manner the sp. gr. of the liquids used in previous experiments.

The common hydrometer.—The scale contained in the tube used in the last experiment might have been graduated to give the sp. gr. by a direct reading. Thus, when floating in water, the scale reading at the surface would be 1·00; and in a liquid of sp. gr. 1·15, the reading on the scale opposite the surface would be 1·15, and so on.

The **common hydrometer** (Fig. 33) is an instrument constructed on this principle. It consists of a graduated glass tube, C, having at one end a large air bulb, A, and a smaller weighted bulb, B. The bulb A floats the instrument, while B serves to keep it upright in the liquid.

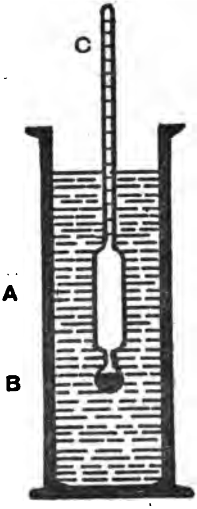

FIG. 33

Expt. 78.—Determine by means of the common hydrometer, the specific gravities of the liquids previously used.

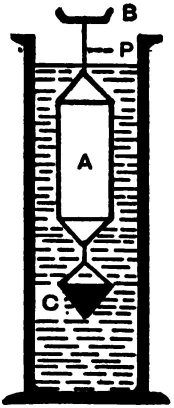

FIG. 34

Nicholson's hydrometer.—This instrument (Fig. 34) can be used for finding the specific gravity either of a solid or of a liquid. It consists of a hollow cylinder, A, connected by a fine stem to a pan, B. There is a mark, P, made on the stem. Attached to the bottom of the cylinder is another pan, C, loaded so as to float the instrument vertically.

Expt. 79.—To determine the sp. gr. of a solid by Nicholson's hydrometer.

Find the weight which must be placed on B to sink the instrument in water until P is just on the surface. Let w = the weight. Place the solid on B, and let w_1 be the weight required on B in addition to sink the instrument to P. Then $w - w_1$ = wt. of solid in air. Now place the solid on the lower pan, C, *leaving the weights w_1 on the upper pan*, B. The instrument now floats at a higher level owing to the upthrust of the water on the solid. Now place w_2 gm. on B to bring P down to the surface. Then w_2 gm. = the upthrust of the water on the solid, *i.e.* the weight of water displaced by the solid.

$$\text{Hence, sp. gr. of solid} = \frac{w - w_1}{w_2}.$$

Expt. 80.—To find the sp. gr. of a liquid, using Nicholson's hydrometer.

Weigh the hydrometer. Find the weights which when placed in B will sink the instrument to P (1) in water, and (2) in the liquid.

Let W = wt. of the hydrometer, w = wt. added to sink it in water, w_1 = wt. added to sink it in the liquid; then

$W + w_1$ = wt. of liquid displaced
$W + w$ = wt. of water displaced

$$\therefore \text{Sp. gr. of the liquid} = \frac{W + w_1}{W + w}.$$

Example 1.—What will a c.dm. of lead (sp. gr. 11·4) weigh in water?

Volume of water displaced, *i.e.* volume of the lead = 1 c.dm. = 1,000 c.c.

Hence, loss of wt. in water, *i.e.* wt. of water displaced = 1,000 gm.

Wt. of lead in air = (1,000 × 11·4) or 11,400 gm.

∴ Wt. of lead in water = 11,400 − 1,000 = 10,400 gm.

Example 2.—A solid floats in mercury (sp. gr. 13·5) with $\frac{2}{3}$ of its volume above the surface. Find the sp. gr. of the solid.

Let V = the volume of the solid, s = its sp. gr., and w = the wt. of unit volume of water; then

Wt. of solid = wt. of mercury displaced.

$$\therefore V \times s \times w = \tfrac{1}{3} V \times 13{\cdot}5 \times w$$

$$\therefore \quad s = \tfrac{1}{3} \times 13{\cdot}5 \text{ or } 4{\cdot}5.$$

ADDITIONAL EXPERIMENTS

81. Find the length of a coil of copper wire without unwinding it.

82. Find the diameter of the bore of a capillary tube in the following manner: Clean and dry the tube, then draw into it a column of mercury about 1 in. long. Move the mercury along the bore and measure its exact length in six different places. Empty the mercury into a weighed watch-glass and weigh it. Density of mercury = 13·59 gm. per c.c. Then calculate the volume of the little cylinder of mercury, and, knowing its length (the mean of the six determinations), obtain its diameter.

83. Place a beaker containing water on the pan of a balance and counterpoise it. Take a solid of known volume V c.c. and suspend it in the water from a stand close by. The pan carrying the beaker is depressed. Why? What weight must be placed on the other pan to restore equilibrium?

EXAMPLES V

1. Find the volume of 100 gm. of mercury: sp. gr. of mercury = 13·6.

2. Two bodies of different bulk weigh the same in water. Which will weigh the more in air?

3. A body whose sp. gr. is 7 weighs 18 gm. in water. What is its weight in air?

4. If a quart of water weighs 40 ozs. and the sp. gr. of petroleum is ·88, find how many gallons of petroleum will weigh 38·5 lbs.

5. A solid soluble in water weighs 17·76 gm. in air, and

8·70 gm. in a liquid of sp. gr. 1·10. What is the sp. gr. of the substance?

6. A solid weighs 56 gm. in air, 35 gm. in water, and 30·8 gm. in a certain liquid. Find (1) the volume of the solid, (2) the sp. gr. of the solid, (3) the sp. gr. of the liquid.

7. A piece of brass weighs 14 gm. in water; when a piece of elm weighing 6 gm. is tied to the brass, the weight of the two in water is 8 gm. Find the sp. gr. of the wood.

8. A body whose volume is 1 c.dm. weighs 1 kgm. in air. What will it weigh in water?

9. Find the sp. gr. of sand from the following weights: Bottle + sand, 40·38 gm. Bottle + sand + water, 91·90 gm. Bottle filled entirely with water, 81·31 gm. Bottle alone, 23·10 gm.

10. A certain body whose volume is 6 c.dm. weighs 15 gm. in air. Find its real weight, having given that the density of air is 0·0013 gm. per c.c.

11. A wooden cylinder of sp. gr. 0·6 and 1 foot in length floats vertically in water. What will be the depth immersed?

12. A piece of wood weighing 30 gm. floats in water with two-thirds of its volume immersed. Find the density and the volume of the wood.

13. What will be the weight of a mixture of 50 c.c. of copper sulphate solution (sp. gr. 1·10) and 30 c.c. of a liquid whose density is 1 gm. per c.c.?

14. A Nicholson's hydrometer weighs 8 oz. The addition of 2 oz. to the upper pan causes it to sink to the mark in one liquid, while 5 oz. are required to produce the like result in another liquid. Compare the specific gravities of the liquids.

15. A body weighing 18 lbs. and of sp. gr. 3 is suspended by a string. What will be the pull or tension of the string when the body is suspended in (1) water, (2) in a liquid of sp. gr. 2?

16. A Nicholson's hydrometer requires 200 grains in the upper pan to sink it to the mark in water; a stone is placed in the upper pan, and the weight required to sink it to the mark is 80 grains; the stone is then placed in the lower pan, and the weight required is 128 grains. Find the sp. gr. of the stone.

CHAPTER IX

THE PRESSURE OF THE AIR

Air is a mixture of several gases, chiefly nitrogen and oxygen. It therefore possesses in a high degree, in common with other gases, the properties of compressibility and expansibility.

It may easily be proved experimentally to possess weight by weighing a glass globe when full of air, and then weighing it after exhausting the air. It is thus found that 1 litre of air weighs about 1·3 gm., and that 1 cu. ft. of air weighs about 1·3 oz.

Pressure of the air.—That air exerts a pressure on a surface with which it is in contact can be shown by various experiments. Thus—

Expt. 84.—Insert a tumbler or beaker, mouth downwards, into water. Notice that the water does not fill the tumbler, showing that the tumbler contains something which prevents the water from filling it. Tilt the tumbler and observe that bubbles escape and rise to the surface, and at the same time more water enters.

Expt. 85.—Fill a gas jar, or a tumbler, with water, lay a piece of writing paper over the top, pressing it down round the edge of the jar, and then invert it (Fig. 35). It will be found that the pressure of the air acting upwards on the paper will support the weight of the paper and of the water in the jar.

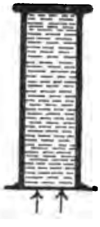

FIG. 35

Expt. 86.—Obtain a cube made of thin tin and having a neck tightly fitted with a cork. Remove the cork and boil a small quantity of water in the vessel. After the air has been driven out and the cube is full of steam, remove the flame and tightly fix in the cork. As the cube cools and the steam inside is condensed, the internal pressure is reduced, and in a short time the cube collapses under the external pressure of the air.

Expt. 87.—Dip one end of a long glass tube into water or mercury. Apply the mouth at the other end and gently withdraw the air from the tube by suction. The liquid rises in the tube in consequence of the pressure of the outer air on its free surface.

The last experiment explains the rise of the liquids in the tubes of Hare's apparatus; also the rise of a liquid in a pipette, a syringe, and a common suction-pump.

Expt. 88.—Cut out a disc of leather and thoroughly soften it in water. Attach a string to the middle of it, and closely fit the leather on a flat stone or on any smooth body. When the string is pulled a vacuum is left underneath, and the pressure of the air holds the edges of the disc against the stone, which can then be lifted.

The following lecture-table experiments require the use of an air-pump :—

Expt. 89.—Place a closed bladder, containing some air, under the receiver of an air-pump; on pumping out the air from the receiver, the pressure of the air on the outer surface of the bladder is diminished, the air inside the bladder expands, and the bladder swells. Readmit the air to the receiver, and the bladder is seen to shrink to its original volume.

Expt. 90.—Close one end of an open glass cylinder by binding on a piece of thin sheet indiarubber. Place the open end of the cylinder on the plate of an air-pump and gradually exhaust the air within it.

The pressure of the air inside the cylinder is thus diminished; the pressure of the outside air forces the elastic covering inwards and finally breaks it.

Expt. 91.—A hollow sphere is formed by two hollow hemispheres (Fig. 36) which fit so closely together as to be air-tight when greased. When the sphere is full of air the hemispheres are easily separated. Exhaust the air from the interior by means of an air-pump; it will then be found that very great force is necessary to separate them, because the pressure of the air on the outer surface is no longer balanced by the pressure inside the sphere. This experiment was first performed by **Otto von Guericke, of Magdeburg**, and is known as the experiment of the **Magdeburg hemispheres.**

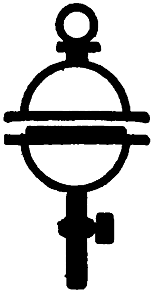

FIG. 36

The above experiments show that the air exerts pressure in consequence of its weight.

A barometer is an instrument for measuring the pressure of the atmosphere.

Torricelli's experiment.—The following experiment was first performed by **Torricelli**, an Italian philosopher, in 1643.

Expt. 92.—Hold in a vertical position a thick-walled glass tube about 34 inches long, ½ inch in diameter, and closed at one end. Fill the tube with mercury, taking care to remove air bubbles; close the open end with a finger and invert the tube. Place the end of the tube and your finger under the surface of some mercury in a basin, and then remove your finger. The mercury falls in the tube a short distance, **BC** (Fig. 37), until the height of the column **AC** above the surface **A** in the basin is about 30 inches, or 760 mm., when the tube is vertical.

The space **BC** above the mercury in the tube contains no air, and is usually called the *Torricellian vacuum.* It is not a perfect vacuum, for it contains a minute quantity of mercury vapour.

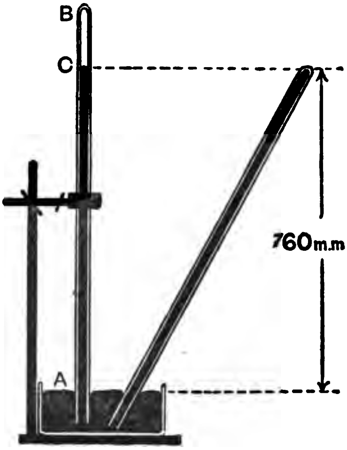

FIG. 37

The column is supported by the pressure of the atmosphere, acting on the surface of the mercury in the basin. This pressure is transmitted through the mercury to the open end of the tube. If the pressure of the atmosphere increases, the mercury will rise in the tube; if it becomes less, then it can only support a smaller weight of mercury, and so the length of the column diminishes.

If a scale be placed alongside the tube for reading off the height of the column, the apparatus constitutes a **simple mercurial barometer.**

The height of the column, **AC**, is called **the height of the barometer** or **the barometric reading.** Thus **the height of the barometer measures the pressure of the atmosphere.**

If the sectional area of the bore of the tube were exactly 1 sq. in., and the length of the column were just 30 ins., the volume of the mercury would be 30 cu. ins., and its weight 14·7 lbs. This mercury column is balanced by a column of air of sectional area 1 sq. in., and whose length extends from the surface of the mercury in the basin to the upper limits of the atmosphere. Therefore **the atmosphere exerts a pressure very nearly equal to 14·7 lbs. wt. per sq. in.** This is called **a pressure of one atmosphere.**

If the tube is slowly tilted to one side, it will be found that the space BC gets less and less, until, when the closed end, B, is less than about 30 ins. above the mercury in the basin, the mercury fills the tube (Fig. 37). But so long as the end B is more than about 30 ins. above A, the vertical height of the column remains the same. Thus **it is the vertical height of the column which measures the pressure of the atmosphere.**

Barometers may be constructed by using water, glycerine, oil, etc., instead of mercury.

The atmosphere supports the weight of a column of mercury 30 ins. in height, and it will support on the same base an equal weight of water. But since mercury is 13·6 times as dense as water, a column of water on the same base, and having the same weight, would extend to a height of 30 × 13·6 inches, or 34 feet; so that **the height of the water barometer is about 34 feet.** Another objection to the use of water is that it evaporates into the vacant space, and thus a good vacuum cannot be obtained. The glycerine barometer does not possess this disadvantage, and is sometimes used.

The heights to which liquids rise in the barometer tubes are inversely proportional to their densities, as is the case with the liquids in the tubes of Hare's apparatus.

The height of a barometer containing a given liquid is the same, whatever the sectional area of the tube, for there is always a column of air of the same sectional area to balance the liquid.

In the simple barometer, if the scale is fixed to the tube, the level of the mercury in the basin should remain the same; but as the mercury in the tube rises, the level in the basin falls, and *vice versâ.* Thus a slight error is introduced in reading the height.

In **Fortin's barometer** this error is corrected by raising or lowering the surface of the mercury in the cistern to the zero of the scale before taking a reading. The bottom of the cistern is made of thin leather, and can be raised or lowered by means of a fine screw which works against it.

The **siphon barometer** (Fig. 38) consists of a U-tube with limbs of unequal length. The shorter limb is open to the atmosphere, while the longer limb, which is about 36 inches in length, is closed. The tube contains mercury, and there is a Torricellian vacuum at the top of the closed limb. When the mercury rises at **A** it falls at **B**. The pressure of the atmosphere, acting at **B**, supports the weight of a column of mercury whose height is *the difference of level of the mercury in the two limbs.*

The pressure of the atmosphere varies at different altitudes. Thus, the height of the mercurial barometer on Mont Blanc is only about 15 inches.

Mercury is used in preference to other liquids because of its great density, so that it requires a shorter tube than water, glycerine, etc. It is easily seen and does not wet the tube, and it is only very slightly volatile. Nevertheless a glycerine barometer is much more sensitive to changes of atmospheric pressure than a mercurial barometer. Glycerine of sp. gr. 1·26 gives a column about 324 in. long, or more than 10 times the height of the mercury column. So that a fall of 1 inch in a mercurial barometer corresponds to a fall of more than 10 inches in a glycerine barometer.

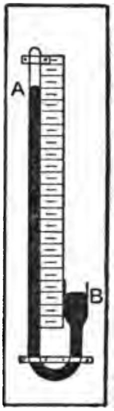

FIG. 38

The **aneroid barometer** consists of a flat, circular box, made of thin corrugated white metal, and exhausted of air. When the pressure increases or decreases, the top falls or rises slightly. This motion is communicated by means of a pillar attached to the top of the box, to a spring which moves a pointer over a dial. The aneroid is chiefly used because it is less cumbrous and fragile than the mercurial barometer.

EXAMPLES VI

1. What is the atmospheric pressure in lbs. per sq. in. when the height of the barometer is 29·5 inches? Sp. gr. of mercury = 13·6.

2. The height of the mercury barometer is 29·5 ins., the sp. gr. of mercury 13·6: find the sp. gr. of oil, when the oil barometer stands at 36·5 feet.

3. What will be the height of the water barometer, when the mercury barometer stands at 29·8 ins.?

4. Find the height of a barometer containing alcohol of sp. gr. ·866 when the height of the mercurial barometer is 29·5 ins.

PART II

PHYSICS

HEAT

CHAPTER X

NATURE OF HEAT—EXPANSION

Nature of heat.—Our senses enable us to distinguish between bodies according as they are hot or cold. If we place a poker in the fire, and after a time remove it, it feels *hot*. A piece of metal which has been exposed to the air on a frosty day feels *cold*. The agent which produces in us these sensations is called *heat*.

According to the old and abandoned theory heat was supposed to be a weightless invisible fluid called *Caloric*, which entered into bodies and caused them to become hotter, whilst bodies grew colder as the Caloric left them. The modern theory supposes heat to be molecular motion—it is a form of energy. The molecules of matter of which a body is composed are in a state of motion, and as the velocity of their motion increases, the substance gets hotter; when the velocity is diminished the body becomes colder.

Temperature.—That a hotter body can impart heat to a cooler body is a matter of observation. For example, a red-hot metal bar is placed in a vessel of water, the bar is cooled, and the water heated; heat is transferred from the bar to the water, the **temperature** of the water is raised, and that of the bar lowered.

Temperature is that condition of a body which determines whether it will communicate heat to, or receive heat from, another body when placed in contact with it. Similarly *level* determines whether one pond will transfer water to, or receive water from, another pond when they are connected by a pipe. But it must be noted that temperature is not heat any more than level is water.

When two bodies are placed in contact, heat flows from the body of higher to the one of lower temperature; if there is no passage of heat between them they are at the same temperature.

In the case mentioned above, on touching the poker which has a higher temperature than the hand, the latter gains heat from the poker and we receive a sensation of warmth; on touching the metal which has been exposed to the air on a frosty day and which has a temperature lower than that of the hand, we receive a sensation of cold, because the hand loses heat to the metal.

Effects of heat.—When heat is applied to bodies it produces various effects, the chief of which may be classified as follows:—

(i) Change of size.
(ii) Change of temperature.
(iii) Change of physical state.

(i) Change of size.—With very few exceptions,[1] all bodies, whether solids, liquids, or gases, increase in size on being heated and contract when cooled. *The amount of expansion is different for different substances;* the expansibility of gases is very much greater than that of liquids, while that of liquids is greater than that of solids. In the case of solids we have—

(*a*) Linear expansion or expansion in length,
(*b*) Superficial expansion or expansion in surface area, and
(*c*) Cubical expansion or expansion in volume,

while we consider only cubical expansion in the case of liquids and gases.

(ii) Change of temperature.—When heat is absorbed by a body it feels hotter—its temperature rises; when it loses heat it feels colder—its temperature falls. The change of temperature which takes place when a given amount of heat is gained or lost varies considerably with different substances.

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[1] Water below 4° C., iodide of silver, stretched indiarubber.

(iii) Change of physical state.—Matter can exist in three states: (1) the solid state, (2) the liquid state, (3) the gaseous state, and many forms of matter may exist in all three states under ordinary conditions. Thus the same form of matter can exist in the solid state as ice, in the liquid state as water, and in the gaseous state as steam. This difference of state is caused by heat. Ice when heated is converted into water, and, if this be supplied with sufficient heat, the water passes into steam. Iron is most familiar to us as a solid, but we know that if it is heated to a very high temperature it melts, and the spectroscope shows us that it exists in the sun as a gas. In fact the addition of heat always tends to convert a solid to the liquid state, and a liquid to the gaseous or vaporous state, whilst the abstraction of heat always tends to change a gas to the liquid state, and a liquid to the solid state.

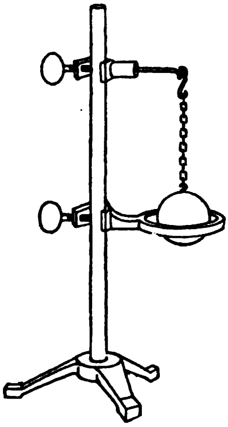


FIG. 39

Expansion of solids.—The expansion produced in solids by an increase of temperature is easily demonstrated by the following experiments :—

Expt. 93.—A ring of metal is made of such a diameter that a metal ball will just pass through it when cold (Fig. 39). Heat the metal sphere in a bunsen flame or in boiling water for a few minutes, and then try to pass it through the ring. It will be found to be too large and will remain on the ring, thus showing that the ball has expanded. Now allow it to cool and notice that after a time it has contracted sufficiently to allow it to pass through.

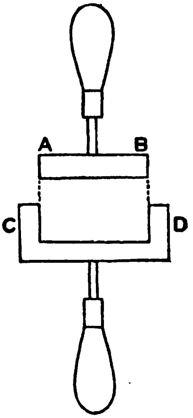


FIG. 40

Expt. 94.—A rod of metal, **AB** (Fig. 40), has its ends accurately squared and is attached to a handle. When cold, the rod fits exactly between the jaws of a metal gauge, **CD**. Remove **AB**, and heat it by placing it in boiling water for a short time. Then try to fit it into the gauge piece. This will be found to be impossible, as the rod is now too long. Hence, the rod **AB** has expanded owing to the heating.

Expt. 95.—A rod of metal, **AB** (Fig. 41), rests in grooves in two metallic supports, **C** and **D**. The end **B** touches the end of an adjusting screw, **E**, which thus prevents expansion towards the right. **E** is turned until the end **A** of the rod presses against the short arm of the lever **P**. Heat the rod by one or two bunsen flames and notice that the end of the pointer gradually moves over the scale **S** towards the right, thus indicating a gradual increase in length of the rod **BA**. Now remove the bunsen burners and observe that the pointer slowly moves back towards its original position as the rod cools.

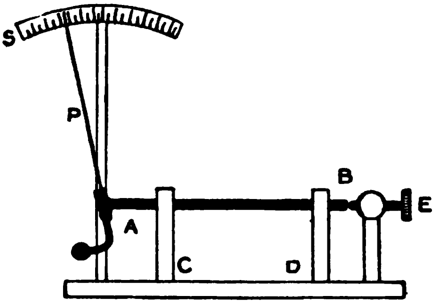


FIG. 41

Expt. 96.—The simple piece of apparatus represented in Fig. 42 gives most satisfactory results. A brass terminal **T** is screwed into a wooden baseboard **XY**. Two small wooden uprights, **W**, **W**, are fixed to the board in the positions shown, and along the top of each is cemented a piece of glass rod. **M** is a small piece of mirror glass, to the back of which is fastened an ordinary sewing needle, **MN**. A piece of thin brass rod is clamped at one end in **T**, the other end resting and pressing on the needle, which is placed across the pieces of glass rod as shown.

Arrange the apparatus so that the rays from a source of light, **S**,—on a bright day the direct rays from the sun serve admirably—fall on the mirror in a direction such as **SM**, and are reflected in the direction **MR**, forming a bright spot on a screen or wall five or six yards away. Then heat the rod **B**; it expands, and, in so doing, gradually rolls the needle and thus turns the mirror through a small angle. The reflected light is consequently deflected, the deflection being indicated by a considerable movement of the spot of light down the screen.

Now remove the source of heat, and observe that the spot of light gradually creeps up the screen, showing that as the rod slowly cools it also contracts.

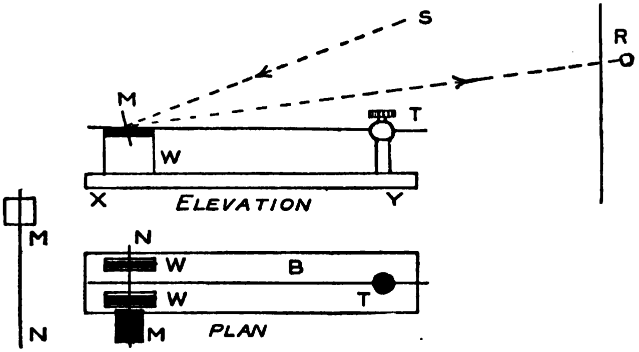


FIG. 42

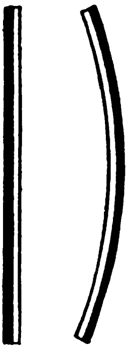
Dimensions of the above piece of apparatus:—

Wooden base, XY	26 cm. × 4 cm. × 1 cm.
Wooden uprights, W, W	3 cm. × 3 cm. × 1 cm.
Mirror, M	1 cm. square.
Metal rod	30 cm. long and 3 mm. diameter.
Brass terminal, T	such as is used in electrical work.

With this apparatus, and the flame of an ordinary wax vesta or taper as the source of heat, the spot of light can be made to move 10 or 12 feet.

Unequal expansion of different solids.

Expt. 97.—Take a straight compound bar made up of a strip of brass and a strip of iron firmly fastened together along their whole length, and heat it in a bunsen flame. Notice that the bar becomes curved, and that the more it is heated the more curved does it become. Each metal expands, but the brass expands more than the iron, and to allow for this the bar becomes curved with the brass on the convex side, thus giving the brass the greater length (ii Fig. 43).



(i) (ii)

FIG. 43

If the above experiment be repeated with compound bars made up of various metals, it will be found that they all become more or less curved, proving that **different metals when heated to the same extent expand by different amounts.** This fact is made use of for practical purposes.

Compensated pendulum.—If the pendulum of a clock is made of one metal only the expansion of the rod in summer causes it to swing more slowly, and thus alters the rate of the clock. Such a clock would lose time in summer and gain time in winter. In **Harrison's gridiron pendulum** (Fig. 44) the length is maintained constant by a suitable combination of metals. It consists of four brass and five steel rods fastened in the cross bars, **A, B, C, D, E,** as shown, the steel rods being represented by the dark lines and the brass rods by the light lines. The central steel rod passes through holes in the lower horizontal cross bars **D** and **E.** The steel rods all expand downwards, and the brass rods upwards. To keep the length of the pendulum constant, *the total upward expansion should be exactly equal to the total downward expansion*, and since brass expands more than steel for the same rise of temperature this result can be obtained by properly adjusting the lengths of the rods.

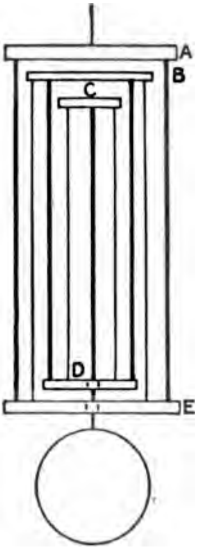


FIG. 44

Force of expansion and contraction.—A metal bar, if prevented from contracting or expanding, exerts enormous force. This may be illustrated as follows :—

Expt. 98.—An iron bar, **AB** (Fig. 45), rests in the notches of a strong cast-iron frame, **DE.** **C** is a short cast-iron rod which passes through a hole at the end **B.** At the other end **A** is a screw thread on which turns a nut, **N.**

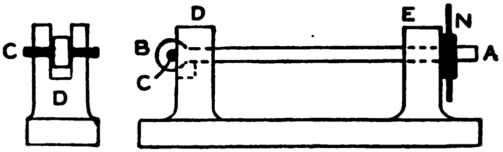


FIG. 45

Heat the bar **AB** in a large bunsen flame, and as it expands screw up the nut so that **C** is kept tightly pressed against the edges of the upright **D**. Then leave the bar to cool; contraction is resisted by the short rod **C**, so that the force exerted on **C** increases and finally becomes so great as to break it.

Useful applications of expansion and contraction.

1. The iron tire of a wheel is slipped on when red-hot. The wheel is then immersed in cold water, and the tire, on contracting, binds together the parts of the wheel.

2. When the walls of a building have bulged outwards they may be restored to their perpendicular position by placing iron rods through opposite walls so as to cross the building. On the projecting ends, outside the building, iron plates are secured by means of nuts. The rods are heated, thus causing them to expand. The plates are then tightly screwed up to the walls by means of the nuts, and the rods allowed to cool, and on contracting the walls are drawn in. This process is repeated till the walls assume their upright positions.

3. In riveting metal plates red-hot rivets are used. A rivet is passed through the holes in the two plates to be fastened together, and is then hammered until the head at each end closely grips the plate. On cooling the plates are drawn together with enormous force.

Practical precautions necessitated by the expansion and contraction of metals.

(*a*) The rails on a railway track are laid with a small space between their ends, to allow for expansion in summer.

(*b*) The bars of a furnace are left free at each end. If they were fixed they would be compelled to bend, when heated, to allow for the expansion.

(*c*) Iron bridges are never fixed at both ends. Usually the girders are each mounted on rollers at the ends, so that they may expand freely. The same applies to iron roof girders of considerable span.

(*d*) For the same reason hot-water pipes are generally laid with telescopic joints, so that they are capable of sliding one within the other.

(*e*) The wooden patterns used for making the sand moulds for metal castings must always be made a little larger than the castings required, to allow for the contraction of the metal on cooling. In the case of cast iron one-eighth of an inch per foot is allowed for this purpose.

(*f*) It will have been noticed that in defining the yard the distance between the two scratches is taken at a particular temperature (62° Fahrenheit). At a higher temperature this distance would be greater than one yard, because of the expansion of the bronze bar.

Expansion of liquids.—Since a liquid must always be contained in a solid vessel, the expansion of the liquid is complicated by that of the vessel which holds it. Thus the amount of expansion observed is not the whole expansion; the expansion actually observed is called the *apparent expansion*. *The real or absolute expansion* is that which would be observed if the vessel did not expand; it is therefore greater than the apparent expansion.

The real expansion = the apparent expansion + the expansion of the vessel.

The expansion of liquids may be easily shown as follows:—

Expt. 99.—Take a long piece of fine glass tubing with a bulb of about 6 cm. diameter blown on one end. Fill the bulb and about one-third of the length of the tube with water coloured with red ink or with a little fluorescein solution. Pass the tube through a long slip of paper as shown in Fig. 46, and make a mark on the paper at the surface of the liquid, as at **A**. Now immerse the bulb in hot water and carefully note that *the surface of the liquid in the tube falls to a point* **B** *for an instant, and then gradually rises* to some such level as **C**. If sufficient heat be supplied the liquid may be made to overflow at the top.

This is due to the fact that the glass bulb is heated first and thus expands first, making more room for the coloured liquid, which then apparently contracts. Soon, however, the liquid is heated, and, having a greater expansibility than glass, it is seen to rise higher and higher in the tube.

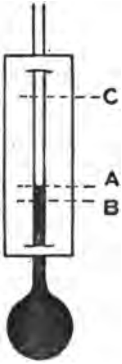


FIG. 46

A 2-oz. round-bottomed flask, fitted with a cork through which passes the end of a long piece of glass tubing of narrow bore, may be used instead of the bulb in the previous experiment.

Unequal expansion of different liquids.—That different liquids have different rates of expansion may be shown by the following experiment:—

Expt. 100.—Fit up three similar and equal glass bulbs or flasks as for the last experiment, and fix them in retort stands with the bulbs in an empty trough. Place methylated spirit in one, coloured water in another, and turpentine in the third, and in each case so adjust the amount of liquid that they stand at the same height in the tubes. Then fill the trough with warm water, and notice that the methylated spirit rises higher than the turpentine or water, and the turpentine higher than the water (Fig. 47).

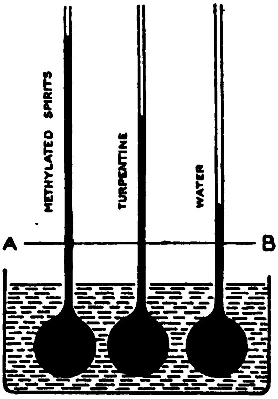


FIG. 47

Irregular expansion of water. Water below 4° C. was mentioned as an exception to the rule that bodies expand on being heated and contract when cooled.

Effect of heat on water.—If heat be applied to ice, which has been cooled to a temperature below 0° C., it follows the usual behaviour of solids and expands. When 0° C. is reached the temperature remains steady till all the ice is melted, during which process **much heat is absorbed or rendered latent.** At the same time every 10 volumes of ice contract to 9 volumes of water.

On continuing the heating, the volume of the water becomes less and less till the temperature reaches 4° C. **At this temperature a given mass of water has its least volume and therefore its maximum density.**

From 4° C. the water gradually expands, as the heating goes on, till the temperature is 100° C., at which temperature the thermometer remains constant till all the water is converted into 1728 times its volume of steam—the pressure being 760 mm.

And, finally, as the heating goes still further, the steam expands at a greater rate, behaving like a gas.

These changes can be expressed in diagrammatic form, as is seen below.

Two straight lines at right angles are taken as axes. Let the vertical line be marked off to represent volumes and the horizontal line to represent temperatures. Ordinates are erected at each temperature to correspond with the volume of the water or ice at that temperature and so a curve can be obtained showing the relation between these quantities.

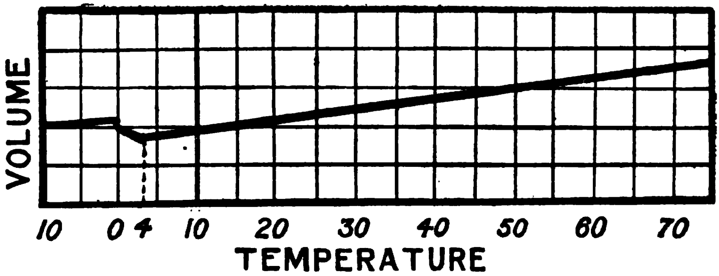


FIG. 48

Expt. 101.—To show the point of maximum density of water.

A tall jar is taken, through the side of which are fixed by means of corks two thermometers. Around its middle is a bath which contains ice (Fig. 49). The jar is filled with water, and the temperatures indicated by the thermometers are read from minute to minute. **The mercury in the lower thermometer gradually sinks to 4° C.**, for as the water cools, as is shown by the upper thermometer, it becomes more dense and sinks to the bottom. Presently there is a uniform temperature of 4° C. As the temperature of the water at the surface falls below this the water now becomes less dense, and consequently does not sink to the bottom, but remains at the top and gets colder and colder till it finally freezes. *Still the lower thermometer indicates* 4° C. This fact is of great importance in nature. For did water steadily contract as it cooled down to zero, then the coldest water would always be found at the bottom, so that ponds, rivers, and seas would freeze there first and then congeal to a solid mass, destroying all their living occupants. And, moreover, the heat of summer would not serve to thaw this accumulation of ice.

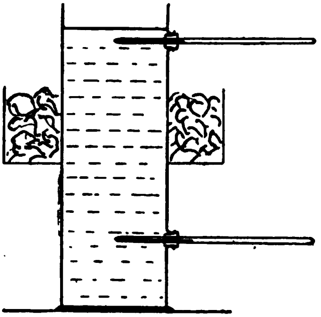


FIG. 49

How liquids are heated.—Liquids do not convey heat in the same way as solids. An iron poker placed in a fire conveys heat by **conduction,** *i.e.* the heat passes from particle to particle of the iron.

But in a liquid the heat is conveyed principally by the liquid itself in **convection currents,** which are due to the difference of density between the hotter and colder parts.

Expt. 102.—To show the convection currents in water.

Place some sawdust in a large beaker full of water. Get the sawdust uniformly diffused through the water. Support the beaker on a stand and heat with a bunsen flame.

As the heating proceeds, notice that the sawdust is carried about with the convection currents of the water. Frequently a complete circulation can be obtained in this way, the sawdust rising up the middle of the beaker, and descending again by the sides.

This circulation of a liquid is also to be observed in Expt. 101, the cooler, denser water sinking to the bottom till a temperature of 4° C. is attained.

Buildings are frequently heated by utilising this property. Water is heated in a boiler in the basement, and then rises through a series of pipes passing through the rooms to a tank placed at the top of the building. As it cools it becomes more dense, and finally at the top of the building it flows down to the boiler by another set of pipes. A constant circulation is thus set up.

Water is a bad conductor of heat.—Heat travels through it very slowly if convection currents are prevented.

Expt. 103.—Load a piece of ice with wire and sink it to the bottom of a test-tube full of water. Heat the water near the top until nearly boiling and notice that the ice does not melt.

Expansion of gases.—Gases, like solids and liquids, expand when heated and contract on cooling.

Expt. 104.—Place before the fire a bladder partly filled with air, and observe that the bladder becomes larger, showing that the heat causes the air inside the bladder to expand. On cooling, the air contracts and the bladder shrinks to its former size.

Expt. 105.—Take a long glass tube with a bulb blown on one end and invert it with the open end dipping in a liquid (Fig. 50, A). Close the hand over the bulb, and thus gently heat the air within. Notice that the air at once begins to expand, and that some of it is forced out of the tube and bubbles through the liquid. Then remove the hand and allow the bulb to cool. The liquid immediately begins to rise in the tube, showing that the air contracts on cooling (Fig. 50, B). If the temperature of the bulb be raised again by the smallest amount, the increase of volume of the air is made evident by an almost immediate fall of the liquid column.

Repeat the experiment, filling the bulb first with coal-gas and afterwards with hydrogen.

Fire-balloons are constructed upon the principle that air expands when heated. When such a balloon is filled with hot air—air which has expanded and which has therefore a density less than that of the surrounding atmosphere—it rises until it reaches a height where the density of the atmosphere is the same as that of the air within the balloon.

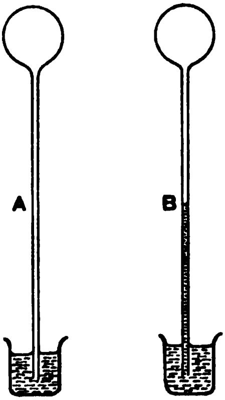


FIG. 50

EXAMPLES VII

1. Explain the difference between temperature and heat. When is the temperature of one body said to be higher than that of another?

2. How is the bulk of a body usually affected by heat? Are there any exceptions to the general rule?

3. Two rulers, the one of iron and the other of brass, are placed one upon the other, and riveted together so as to form a single ruler of two layers. At a temperature of 60° the compound ruler is straight. Describe and explain what occurs when it is successively exposed to a temperature of 30° and to a temperature of 90°.

4. Make a diagram of the gridiron pendulum, and state fully the principle on which its construction depends.

5. A bladder, filled tightly with air, is carried from a very hot to a very cold room. What will be the effect of the low temperature on the air inside the bladder?

6. How would you prove that zinc expands more than copper when rods of the two metals are heated through the same range of temperature?

7. Explain why a fire-balloon ascends in the air, and why at some altitude it must stop rising.

8. What is the usual effect of heat on the volume of a substance? Mention an exception in the case of water, and a natural phenomenon that this accounts for.

9. Water is said to have its maximum density at 4° C. Explain what this means.

In what respect is the behaviour of mercury different from that of water when both are gradually warmed from 0° C?

10. Distinguish between the absolute and the apparent expansion of a liquid.

11. What practical precautions are necessitated by the expansion and contraction of metals?

12. Mention some useful applications of the expansion and contraction of metals.

CHAPTER XI

THE PRINCIPLE AND CONSTRUCTION OF THE THERMOMETER

THE **thermometer** is an instrument used to measure temperature. When heat is applied to bodies it produces, among other effects, change of volume. The principle of the thermometer depends upon the fact that the general effect of raising the temperature of the body is to cause it to expand, and the expansion of liquids is usually made use of in the construction of these instruments.

A gas thermometer, containing either hydrogen, nitrogen, or air, is often employed for special purposes. Such a thermometer gives accurate readings of temperature over a great range; but its form is not convenient for ordinary use.

Ordinary thermometers contain either mercury or coloured alcohol.

Mercury is the best liquid to use for several reasons:—

(1) Its boiling point is high (350° C.), and its freezing point low (–40° C.).

(2) It takes very little heat from the substance whose temperature it measures.

(3) It is a good conductor of heat, and therefore it quickly attains the temperature of a substance in contact with it.

(4) Its increase of volume for 1° rise of temperature is nearly uniform, and is large in amount.

(5) It does not wet the tube and is easily seen.

Alcohol (freezing point about –130° C.) is used in thermometers which are intended to register low temperatures.

Construction of a thermometer.

Expt. 106.—To fill an alcohol thermometer.

Blow a strong bulb about ¾ inch in diameter at the end of a piece of thermometer tubing 10 inches long. Join a small glass funnel to the open end by means of a short piece of rubber tube, and clamp the tube in a vertical position (Fig. 51). Place some coloured alcohol in the funnel, and allow the flame of a bunsen burner to gently heat the bulb. The air inside the latter will expand and will be seen to rise in bubbles through the alcohol in the funnel. Now remove the flame, and as the bulb cools the air remaining in it contracts, and some of the alcohol will be forced into the bulb by atmospheric pressure. Again heat the bulb, sufficiently to boil away the alcohol contained in it, and then again allow the bulb to cool. The bulb and tube ought now to be filled with alcohol. Remove the funnel, and observe that as the bulb cools to the ordinary temperature the alcohol contracts.

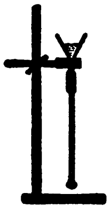


FIG. 51

Then find the range of the tube by immersing the bulb in water and raising its temperature about 20°; it should be about 1 inch for 10°; if it is not about this range another bulb must be blown.

At a point about 8 inches from the bulb, draw out the tube to a narrow neck, and then place the bulb and part of the tube into water at a temperature a little higher than the highest—say 60° C.—which it is intended to register. Some alcohol will overflow. Then whilst the tube is full, quickly seal it off at the neck with a small blow-pipe flame, and after removing the bulb from the water, complete the sealing so as to strengthen the end.

The fixed points of a mercurial thermometer.—A thermometer tube may be filled with mercury in a similar manner. We then need two fixed points in order to form a scale of temperature. The two temperatures adopted are those of melting ice, and of the steam given off by water boiling under an atmospheric pressure of 760 mm. of mercury.

The **freezing point**, as the lower fixed point is called, is determined by placing the bulb and lower part of the stem in a vessel—often a large glass funnel—containing pounded ice which has been carefully washed. After standing for about fifteen minutes the level of the mercury becomes stationary, and, with a file, a mark is carefully scratched on the glass at this point.

The **boiling point** is determined by placing the thermometer in the **steam** given off by water boiling under *standard atmospheric pressure.* When the level of the mercury becomes stationary a mark is scratched on the glass as before.

If the pressure at the time be not that due to 760 mm. of mercury, the mark will give the boiling point of water under the pressure existing at that time. In such a case a correction is required.

Graduation of the thermometer.—The part of the stem between these two fixed points is divided into a certain number of equal parts called *degrees.* Three scales are in general use.

Name of scale.	Freezing point marked.	Boiling point marked.	Number of degrees or divisions between fixed points.
Centigrade (C.)	0°	100°	100
Fahrenheit (F.)	32°	212°	180
Réaumur (R.)	0°	80°	80

Thus the freezing point is 0° C., 32° F., and 0° R., whilst the boiling point is 100° C., 212° F., and 80° R.

It follows that

100 Cent. degrees = 180 Fah. degrees = 80 Réaumur degrees,
or 1 Cent. degree = $\frac{9}{5}$ Fah. degree = $\frac{4}{5}$ Réaumur degree.

The degrees may be continued above and below the fixed points.

Let Fig. 52 represent a thermometer to which all three scales are attached, **A** being the freezing point and **B** the boiling point, and suppose the top of the mercury column to be at **D**. Let C., F., R., be the respective readings on the three scales for this temperature. Now the distance **AD** represents C. Centigrade degrees, F. − 32 Fahrenheit degrees, and R. Réaumur degrees, whilst **AB** represents 100 Centigrade, 180 Fahrenheit, and 80 Réaumur degrees respectively. The fraction $\frac{\mathbf{AD}}{\mathbf{AB}}$ being the same on each scale, we have—

$$\frac{\text{C.}}{100} = \frac{\text{F.} - 32}{180} = \frac{\text{R.}}{80}.$$

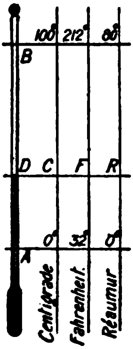


FIG. 52

From these equations, when the reading on one scale is given, the corresponding readings on the other scales can be found.

Example 1.—What is the Centigrade reading corresponding to 28° F.?

$$\frac{\text{C.}}{100} = \frac{\text{F.} - 32}{180}; \therefore \frac{\text{C.}}{100} = \frac{28 - 32}{180}; \therefore \text{C.} = -2^\circ\tfrac{2}{9}.$$

Example 2.—Convert 15° C. to Fahrenheit scale reading.

$$\text{We have } \frac{15}{100} = \frac{\text{F.} - 32}{180}; \therefore \text{F.} = 59^\circ.$$

EXAMPLES VIII

1. Express in the Centigrade scale the following Fahrenheit temperatures: −40°, 0°, and 180°.

2. Express in the Fahrenheit scale the following Centigrade readings: 4°, 40°, and −10°.

3. What are the readings on the Fahrenheit and Centigrade scales corresponding to 60° R.?

Expt. 107.—Graduation of the alcohol thermometer.

The alcohol thermometer is graduated by comparison with a Centigrade mercurial thermometer. Mark the freezing point in the manner explained above. Then place the alcohol and mercury thermometers in a beaker of ice-cold water. Very slowly heat the water and keep it well stirred. Quickly make fine file marks on the tube where the alcohol surface stands, as the mercury thermometer reads 5°, 10°, 15°, 20° . . . 60°. Lay the thermometer upon a small block of wood, 10 inches by 2 inches by ½ inch, having paper pasted upon one face, and transfer the marks to the paper. Then remove the thermometer and carefully rule the degree marks upon the paper. Observe that the distances corresponding to one degree are greater the higher the temperature.

The thermometer may then be fixed to the board, either with copper wire passing through holes in the board, or by two bent strips of metal passing over the tube and fixed into the wood.

EXAMPLES IX

1. Describe the method of filling and graduating an alcohol thermometer.

2. Describe the method of determining the fixed points of a mercury thermometer, and mention any precautions that ought to be observed.

3. What are the advantages and disadvantages of alcohol as compared with mercury for use in a thermometer? Why is water an unsuitable liquid to use?

4. What do we mean by temperature of a substance? State what property of bodies we employ for the purpose of indicating temperatures.

5. Define a degree centigrade.

Explain carefully why an alcohol thermometer is graduated by direct comparison with a mercury one.

CHAPTER XII

EXPANSION OF GASES

Change of volume of gases by heat.—The law of the expansion of gases when heated is known either as **Charles', Dalton's, or Gay-Lussac's Law.** It is said to have been first discovered, though not published, by Charles in 1786.

The law states that **the volume of a given mass of any gas which is kept at a constant pressure, increases for each rise of temperature of 1° C. by a definite fraction ($\frac{1}{273}$) of its volume at 0° C.**

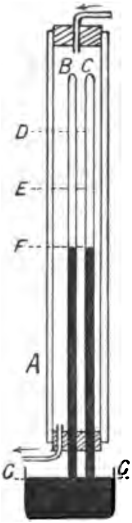


FIG. 53

If the temperatures be stated in degrees Fahrenheit the fraction is $\frac{1}{491}$.

The ratio of the *increase* in volume of a given mass of any gas for a rise of 1° C., to the volume at 0° C., is called **the co-efficient of cubical expansion** of the gas. It follows that for any gas the value of this co-efficient is $\frac{1}{273}$ or ·00366. Hence, **equal volumes of all gases under the same pressure expand by the same amount for a given rise of temperature.** This may be verified as follows:—

Expt. 108.—Arrange two similar barometer tubes fitted up as simple barometers side by side, and as shown in Fig. 53. **A** is a wide tube which acts as a steam jacket. Suppose **DG** to be the height of the barometers. Pass air up the tube **B** and coal-gas up the tube **C** until the mercury falls to the same level **E** in each tube. We have then equal volumes and equal lengths of the two gases at the same temperature. Then allow steam to enter the tube **A** at the top and escape at the bottom. The gases are heated through the same range of temperature and expand by the same amount, moving down to the same level **F**.

Let the volume of a given mass of gas at 0° C. be v_0 cubic centimetres. Then the volume at 1° C. is $v_0+\frac{1}{273}v_0$, at 2° C. it is $v_0+\frac{2}{273}v_0$, and at t° C. it is $v_0+\frac{t}{273}v_0$.

Hence, if v_t cubic centimetres be the volume at t° C.

$$v_t=v_0+\frac{t}{273}v_0=v_0\left(1+\frac{t}{273}\right).$$

Also if v_T be the volume at T° C.

$$v_T=v_0+\frac{T}{273}v_0=v_0\left(1+\frac{T}{273}\right).$$

Therefore by division we get

$$\frac{v_T}{v_t}=\frac{\left(1+\frac{T}{273}\right)}{\left(1+\frac{t}{273}\right)}=\frac{273+T}{273+t}$$

Absolute Temperature.—The volume of a gas kept at constant pressure decreases $\frac{1}{273}$ of its volume at 0° C. for every 1° C. fall in temperature. Then, if we could suppose a gas to continue to contract according to this law for any number of degrees fall of temperature, at −273° C. its volume would become *zero*. This temperature of −273° C. is called the **absolute zero**, and temperatures measured from this zero are called **absolute temperatures.**

Then −273° C. = 0° Absolute, 0° C. = 273° Absolute, and t° C. = $(273+t)$° Absolute.

The above equation may then be written—

$$\frac{v_T}{v_t}=\frac{273+T}{273+t}=\frac{\text{Absolute temperature corresponding to } T^\circ \text{ C.}}{\text{Absolute temperature corresponding to } t^\circ \text{ C.}}.$$

In other words, **the volume of a given mass of gas at constant pressure is proportional to its absolute temperature.**

This form of Charles' Law is the most useful for numerical calculation.

The volume of a given mass of gas may be altered considerably by a small change of pressure, even though the temperature be kept constant.

The relation between the pressure and volume when the temperature remains constant was first stated by the **Hon. Robert Boyle in 1662**, and is known as **Boyle's Law.**

Boyle's Law.—If the temperature be constant, the volume of a given mass of gas varies inversely as the pressure to which it is subjected.

In other words, if the pressure on a gas be doubled, its volume is halved; if the pressure be halved, its volume is doubled, and so on.

Let p denote the pressure of the gas when the volume is v, then if the pressure be changed to $n \times p$, the volume becomes $\frac{v}{n}$. Now the product of the original pressure and volume is $p \times v$. Similarly, the product of the new pressure and volume is $np \times \frac{v}{n}$, or $p \times v$; that is—

original pressure × original volume = new pressure × new volume;
or, $p \times v = p_1 \times v_1$.

Thus **the product pressure × volume remains constant** in value, for one quantity decreases exactly in proportion as the other increases.

Expt. 109.—To verify Boyle's Law.

The apparatus consists of a strong glass tube, AB (Fig. 54, frontispiece), closed at its upper end A, and connected at its lower end, by means of a piece of rubber tube, with another glass tube, C. The latter tube is open at the top and is fixed to a wooden block, which can be clamped, by means of a screw at the back, in any position on the vertical slot shown in the upright stand. The rubber tube and lower part of the glass tubes contain mercury, whilst the upper part of the tube AB is filled with dry air.

This enclosed air is the given mass of gas on which the experiment is to be performed.

The level of the mercury in each tube is read on the scale by the aid of an adjustable pointer. The volume of the air is proportional to the length of the part of the tube which it occupies. Thus the volume may be represented by so many cm. of the tube. This length is reckoned from a point, A, a little lower than the top of the tube, to allow for the rounding off where the tube is closed.

The pressure on the enclosed air is varied by raising or lowering the sliding tube.

Let the height of A above the base D, as measured on the scale, be h_1.

Let **B** and **C** be the surfaces of the mercury in the tubes **A** and **C** respectively; and let the height of **B** above **D** be h_2, and of **C** above **D** be h_3.

Adjust the sliding tube until the mercury stands at the same level in the two tubes. The air is then at atmospheric pressure. Determine this pressure by reading the height, h, of the barometer.

Then in this particular case the product of the pressure and volume is $h\times(h_1-h_2)$, or $h\times(h_1-h_3)$, for $h_2=h_3$. Now raise the sliding tube. The mercury in **AB** rises, only not so quickly, and thus diminishes the volume of the air. Read off the new values of h_2 and h_3. The new volume is h_1-h_2, and the pressure $h+(h_3-h_2)$, for the pressure is increased by the column of mercury between the two levels. Find the value of the product $[h+(h_3-h_2)]\times(h_1-h_2)$.

To make measurements with pressure less than atmospheric, the sliding tube must be moved below the position where $h_2=h_3$. In such a position the pressure on the enclosed air is $h-(h_2-h_3)$. Enter your results as in the following example :—

h_1	h_2	h_3	Pressures.			Volume. h_1-h_2	$P\times V$
			Atmosphere.	Mercury.	Total.		
59·7	47·9	47·9	75·4	0	75·4	11·8	889
59·7	49·3	59·0	75·4	+ 9·7	85·1	10·4	885
59·7	49·9	64·6	75·4	+ 14·7	90·1	9·8	883
59·7	47·1	42·3	75·4	– 4·8	70·6	12·6	889
59·7	44·0	25·7	75·4	– 18·3	57·1	15·7	896
59·7	43·2	22·1	75·4	– 21·1	54·3	16·5	895

It will be found that the products of pressure and volume are constant within the limits of experimental error.

Plot your values of pressure and volume on squared paper, taking the volumes as abscissæ. The curve should be what is known as a *rectangular hyperbola.*

Careful investigations have shown that in no case does any gas strictly obey Boyle's Law. However, for gases which are very difficult to liquefy, such as air, oxygen, nitrogen, and hydrogen, the deviation from the law is extremely small. Most gases are a little more compressible than Boyle's Law would lead us to believe. Hydrogen is an exception, being rather less compressible than the law would imply.

On the Continent Boyle's Law is known as *Marriotte's Law.*

The law connecting the pressure, volume, and temperature of a gas, when all three quantities are allowed to vary, is obtained by combining Boyle's Law and Charles' Law, as follows :—

Let p_1, v_1, t_1 represent the pressure, volume and temperature of a given mass of gas, and suppose these quantities are allowed to vary until p_2, v_2, t_2 represent the new pressure, volume and temperature of the mass of gas. Suppose the changes to take place in two stages.

(1) Change p_1 to p_2 whilst t_1 remains constant.

Then v_1 changes to v say, and by Boyle's Law we have—

$$p_1v_1 = p_2v, \text{ or } v = \frac{p_1v_1}{p_2}.$$

(2) Then change t_1 to t_2, keeping the new pressure (p_2) constant. The volume v then changes to v_2, and by Charles' Law

$$\frac{v}{v_2} = \frac{273 + t_1}{273 + t_2}, \text{ but } v = \frac{p_1v_1}{p_2};$$

$$\therefore \frac{p_1v_1}{p_2v_2} = \frac{273 + t_1}{273 + t_2}.$$

Or, if T_1, T_2 be the absolute temperatures corresponding to $t_1°$ C. and $t_2°$ C., we can write the last equation thus—

$$\frac{p_1v_1}{p_2v_2} = \frac{T_1}{T_2}, \text{ or } \frac{p_1v_1}{T_1} = \frac{p_2v_2}{T_2}.$$

EXAMPLES X

1. A litre of air is measured under a barometric pressure of 740 mm. of mercury. What volume will it occupy under a pressure of 760 mm.?

2. 100 c.c. of oxygen are measured at 100° C. What volume will the gas occupy at $-100°$ C.?

3. A litre of hydrogen is measured at 0° C. and 760 mm. pressure. Find its volume at 15°·5 C. and 740 mm. pressure.

CHAPTER XIII

CHANGE OF STATE

Fusion and Vaporization.

Expt. 110.—Nearly fill a beaker with pounded ice and insert the bulb of a thermometer. Slowly heat the beaker and gently stir the contents. Notice that *the thermometer remains at* 0° *C. till all the ice has melted.*

The change from the solid state to the liquid state is called **melting** or **fusion**, and the temperature at which it takes place is called the **melting point** or **fusing point** of the substance. The ice melts or fuses at 0° C. The reverse process is called **solidification** or **freezing.**

Then the temperature gradually rises and minute bubbles collect on the sides of the beaker. Notice that these, after a short time, rise *gently* to the surface and escape. They are formed from the air which was dissolved in the water. (See Expt. 48, p. 138.)

As the temperature rises from 0° C. to 100° C. water is continually passing off into *invisible* vapour by the process of *evaporation.* This vapour may be condensed in extremely minute particles by the cold air to form a *visible mist or cloud* just above the surface of the water.

Evaporation is the process by which a liquid gradually changes into vapour at its free surface.

Evaporation may go on at all temperatures and increases rapidly as the temperature rises, the vapour being formed *only at the surface.*

Soon, bubbles of *vapour* are formed at points in the liquid near the source of heat; these rise and are condensed in the upper colder layers of the liquid, the condensation producing the familiar sound known as "*singing.*" At length bubbles of vapour rise to the surface and break, and the process of **boiling** or **ebullition** has begun. Notice that the thermometer then reads 100° C. Allow the water to boil for some time, and observe that *the temperature ceases to rise,* and remains at 100° C. until all the water has boiled away.

Ebullition or boiling is the rapid conversion of a liquid into vapour, the vapour being formed not at the surface but within the liquid itself.

The temperature at which ebullition occurs is called the **boiling point** of the liquid.

Latent heat.—The preceding experiment showed that the temperature remained at 0° C., whilst the change from the solid to the liquid state was taking place; that is, until the last particle of ice had been melted; and yet during that change heat was being supplied to the beaker and contents. Also whilst the water was being converted into steam, the thermometer remained at 100° C., although heat was being absorbed. The heat which is thus absorbed *without raising the temperature of a body*, is employed in changing its state, from a solid to a liquid, or from a liquid to a gas, and is called **latent heat.** During the condensation of a gas and the solidification of a liquid, the latent heat is liberated.

The amount or quantity of heat which passes from one body to another may be measured in terms of a suitable unit.

The **unit of heat**, or unit quantity of heat, **is the amount of heat required to raise the temperature of 1 gm. of water through 1° C.**

The latent heat of fusion of ice is the number of units of heat absorbed by 1 gm. of ice at 0° C. in changing into water at 0° C. Its value is large—about 80; that is to say, to change 1 gm. of ice at 0° C. into water at 0° C. requires 80 times as much heat as is required to raise the temperature of 1 gm. of water 1° C. *The same amount of heat is given out when 1 gm. of water at 0° C. is converted into ice at the same temperature.*

The latent heat of vaporization of water, or the latent heat of steam, **is the number of units of heat absorbed by 1 gm. of water at 100° C. in changing into steam at 100° C.** Its value is also very great—about 537; that is to say, to convert 1 gm. of water at 100° C. into steam at 100° C., 537 times as much heat is required as will raise the temperature of 1 gm. of water through 1° C. *The same amount of heat is given out by 1 gm. of steam at 100° C. when condensed to water at 100° C.*

Condensation is the process by which a vapour is changed into liquid.

The process of turning a liquid into its vapour, and then condensing the vapour, is called distillation.

Expt. 111.—Distillation of water.

Fit up a retort in a stand, as in Fig. 55. Place on the neck of the retort a small flask to act as receiver. Keep the receiver cool in a trough of water. Place water in the retort till it is about one-third full, and fit the tubulure of the retort with a cork.

Heat the retort gently. The temperature rises to about 100° C., and then the water boils. The evolved steam begins to condense on the cold neck of the retort and runs down into the receiver. This distilled liquid is usually termed the **distillate**.

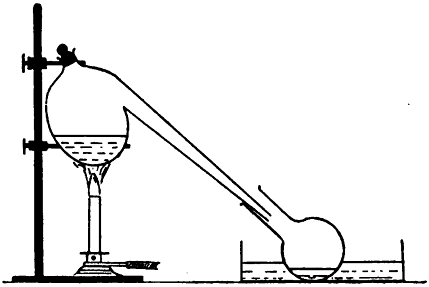


FIG. 55

Cold produced by evaporation.—It has just been shown that whenever a liquid passes to the state of vapour a considerable amount of heat is absorbed. If no external source of heat is used, the heat that is thus rendered latent is obtained from the remaining liquid or from any body in contact with it. Hence *the evaporation of a liquid tends to cool the liquid and also surrounding objects.* This cooling effect is very noticeable in the case of ether.

Expt. 112.—Pour a little ether on a piece of muslin or cotton wool wrapped round the bulb of a thermometer, and observe that its rapid evaporation produces an appreciable fall of the mercury column.

Expt. 113.—Place a few drops of ether on the hand. The liquid quickly evaporates, and the heat required for evaporation is taken from the hand, which consequently feels colder.

Expt. 114.—Place a small quantity of water in a test-tube, and place the latter in a beaker containing ether. Stand the beaker in a film of water which has been placed on a wooden block. Then blow a stream of air through the ether by means of a bellows. The ether evaporates very rapidly, and consequently a large amount of heat is absorbed. Notice that the film of water freezes and fixes the beaker to the block, and also that in a few minutes ice begins to form in the test-tube.

In India and other hot countries water is kept cool by placing it in porous earthenware jars. The water passes through the pores to the outer surface, where evaporation takes place. The heat thus absorbed is taken from the vessel and from the water, which is thus kept cool. Canvas bags are sometimes used for a similar purpose.

The evaporation of the moisture in damp clothing reduces the temperature of the body.

Freezing mixtures.—Heat is absorbed in changing a solid into liquid. When snow or broken ice is mixed with common salt, the ice melts and the salt dissolves; heat is absorbed during the liquefaction and the mixture cools. The resulting temperature may be as low as − 22° C. Such a mixture is called a **freezing mixture.**

Expt. 115.—Effect of dissolved salt on the freezing point of water.

Place some snow or some ice-shavings in a beaker and mix with it about one quarter of its bulk of common salt. Then hold in the mixture the bulb of a thermometer graduated below 0° C., and note the lowest temperature registered. It will be found to be about 14 or 15 degrees below the zero. Thus ice and salt when mixed form a **freezing mixture.**

Now place a test-tube about half filled with water into the mixture. Put the bulb of the thermometer into the water, and notice that the temperature falls to 0° C., and that ice begins to form *at the same temperature.* Then remove the thermometer and test-tube. Melt the ice in the test-tube and add a little salt to the water. Again stand the tube in the freezing mixture. It will be found that ice does not form until a temperature *below* 0° C. is reached. Record the temperature.

Expt. 116.—Find the temperature produced when powdered sodium sulphate, ammonium nitrate or ammonium chloride is dissolved in water.

Also mix pounded ice and crystallised calcium chloride and observe the temperature produced.

The above experiments show that **the freezing point of water is lowered when salts are present in solution.**

Expt. 117.—To determine the melting points of such substances as paraffin wax, lard, beeswax, tallow, etc.

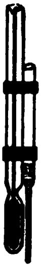


FIG. 56

First method.—Melt some of the substance in a beaker and immerse the bulb of a thermometer in the liquid. Remove the thermometer and notice that it is covered with a thin transparent layer of the liquid. Allow the thermometer to cool, and observe that as this layer solidifies it assumes a frosted appearance. The reading on the thermometer when this takes place is the melting point of the substance. Now place the bulb, with the solidified substance on it, in a beaker of water and gently heat the water. The temperature at which the layer becomes transparent is the melting point of the substance. Take the mean of the two results obtained in the above manner.

Second method.—Soften a piece of ordinary glass tubing in a bunsen flame, and then draw it out to form a capillary tube.

Nearly fill the tube with the melted substance, and then allow it to cool. Attach the tube to the bulb of a thermometer by means of two rings cut from a piece of rubber tubing (Fig. 56). Place the bulb in water contained in a beaker. Heat the water and read the thermometer at the moment that the substance becomes transparent. Now let the water cool, and note the temperature at which the substance becomes opaque. The mean of the two observations gives the melting point of the substance.

Change of volume accompanying change of state.—As a general rule substances expand when they melt, so that the liquid is less dense than the solid at the same temperature.

Expt. 118.—Melt some paraffin wax in a test-tube; then drop in a few small pieces of the solid paraffin and observe that they sink; therefore paraffin expands on melting.

Next, allow the liquid wax to solidify, and then notice that the upper surface is depressed in the centre, showing that paraffin contracts very much on solidifying.

Gold and silver expand on melting and contract on solidifying, so that it is impossible to get sharp castings with these metals. For this reason gold and silver coins and medals are stamped.

A few substances, however, contract when they melt and expand during solidification, so that the solid floats in the liquid of the same temperature; for example, water expands on freezing; consequently ice floats in water.

This explains **the bursting of water-pipes** in winter. During a frost the water in the pipe freezes and expands, bursting the pipe, and filling the rent with ice, so that the damage is not at once made evident. When the thaw comes, the ice melts, and the water is allowed to escape.

In a similar manner frost assists in the disintegration of rocks; it helps the farmer by breaking up the rough lumps left by the plough, and it damages the surfaces of roads and buildings.

Cast-iron and brass, also, expand when passing from the liquid into the solid state, so that a piece of the solid floats on the molten metal.

These metals are well adapted for making sharp castings, since the expansion on solidification forces the metal into every interstice of the mould. (*Solid* cast-iron and *solid* brass *contract on cooling.*)

Influence of pressure on the melting point.—A substance which contracts on melting has its melting point lowered by an increase of pressure.

Ice has its melting point lowered by pressure.

A substance which expands on melting, like wax, has its melting point raised by pressure.

If ice be forcibly compressed partial liquefaction takes place. Ice-cold water (0° C.) does not solidify if it is prevented from expanding, and if the pressure applied be sufficiently great it can be kept in a liquid state at a temperature much below 0° C. **The freezing point of water** (or the melting point of ice) **is thus lowered by an application of pressure.**

Expt. 119.—Press together two pieces of ice; the ice surfaces in contact melt, and when the pressure is relieved the water thus formed freezes again and unites the pieces.

This freezing again is called **regelation.**

Pieces of ice floating in warm water can be pressed together and frozen into one piece.

The following experiment also illustrates the phenomenon of regelation:—

Expt. 120.—Support a block of ice at each end; pass a loop of copper wire round it between the supports and then attach a heavy weight to the wire underneath. Notice that the wire gradually cuts its way through the block *without actually dividing it into two parts.*

The ice just below the wire is melted by the pressure, and the water thus formed flows round the wire and freezes above it.

Boiling point.—The temperature at which a liquid boils depends upon the pressure on the surface of the liquid.

The boiling point of water is 100° C. *only when the atmospheric pressure is that due to 760 mm. of mercury.* The following table, by Regnault, gives the boiling point of water under various atmospheric pressures:—

Boiling points (degrees Centigrade).	Pressure (mm. of mercury).	Boiling points (degrees Centigrade).	Pressure (mm. of mercury).
99·0	733·2	99·8	754·6
99·2	738·5	100·0	760·0
99·4	743·8	100·2	765·5
99·6	749·2	100·4	771·0

Thus, under an atmospheric pressure greater than 760 mm. of mercury water boils at a higher temperature than 100° C., and under a pressure less than 760 mm. at a temperature less than 100° C.

Expt. 121.—To show the effect of reduction of pressure on the boiling point of water.

Boil some water in a flask, which can be closed with a well-fitting cork. Steam is evolved, which, after a time, drives out most of the air from the flask. Remove the flame and cork the flask tightly. Cool by holding the flask under a stream of water. Notice that the water begins to boil, although its temperature is under 100° C. The atmosphere of steam which was above it has now been condensed, thus a great reduction of pressure has taken place, resulting in the water boiling at a lower temperature.

In a steam boiler, where the pressure is very great, the boiling point of water is raised considerably. In high altitudes, where the atmospheric pressure is small, cooking by means of boiling water in the ordinary way is impossible, because the water is not sufficiently hot. Vessels are used, therefore, which can be screwed up, and in which the pressure of the steam can be raised high enough to enable the requisite temperature to be attained.

Expt. 122.—Effect of salt on the boiling point of water.

Fit up the apparatus illustrated in Fig. 57, the flask being about half full of water. The thermometer must be graduated above 100° C.

To prevent boiling with bumping, place one or two small pieces of broken glass in the flask. Boil the water, and observe that the reading is 100° C. when the bulb is in the steam. Remove the thermometer and dry the bulb. Add a little salt to the water and again boil. Replace the cork so that the thermometer bulb is in the steam given off by the boiling solution, and observe the temperature. It will be found to be the same as before (100° C.). Now lower the thermometer bulb into the boiling solution, and observe that the temperature is above 100° C. Record the temperature.

Thus **the effect of dissolving salt in water is to raise the boiling point**, and the more salt is dissolved the higher is the boiling point, until the water becomes saturated, when the boiling point is constant.

The steam given off from such a solution is pure water vapour.

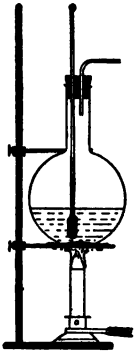


FIG. 57

Note.—To determine the boiling point of a salt solution the thermometer bulb must be placed *in the solution.*

Expt. 123.—To determine the boiling point of a liquid.

Fig. 58 illustrates the apparatus used. A large boiling tube 1·5 inches in diameter is fitted with a cork, through which passes a test-tube 1 inch in diameter. The latter has a hole fused through it at **A**, and is also fitted with a cork, through which passes a thermometer and a piece of bent glass tube.

Put some liquid in the larger tube, refix the cork and test-tube and boil the liquid. The vapour surrounds the test-tube and enters it at **A**, passing to the bottom and then out through the escape tube. Note the temperature when the level of the mercury becomes stationary. This is the boiling point under the pressure indicated by the barometer. Record the temperature and the height of the barometer.

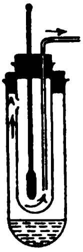


FIG. 58

Find in the above manner the boiling points of alcohol and turpentine.

Sublimation.—There are some solids which change directly into vapour without passing through the intermediate liquid state. This process is termed **sublimation**. If a vapour thus formed is allowed to condense on a cold surface, a thin layer of the solid, called a **sublimate**, is formed.

Expt. 124.— Heat two or three crystals of iodine in a test-tube. Notice that the iodine does not melt, but passes off into a deep purple vapour, *i.e.* it sublimes. Also, some of the vapour condenses again to the solid state on the cooler parts of the tube.

Repeat the above, using sal-ammoniac instead of iodine.

Arsenic, camphor, and solid carbon dioxide also evaporate without liquefaction and condense again upon colder surfaces.

During a long frost snow evaporates and slowly disappears at temperatures below the freezing point. This is a familiar instance of the direct evaporation of a solid.

EXAMPLES XI

1. Distinguish between evaporation and ebullition, and state the conditions upon which the *rate* of evaporation of a liquid depends.

2. Describe the changes which take place in a piece of ice originally at – 3° C., when heat is applied to it so as to raise the temperature to 105° C.

3. A glass bottle and a bottle of porous earthenware are both filled with water and exposed to the air side by side. Usually the water in the earthenware bottle becomes decidedly colder than that in the glass. Why is this?

4. In determining the boiling point of water, why must the barometer be consulted? Describe any experiment to show that water may be made to boil at a temperature lower than 100° C.

5. What is meant by the "boiling point" of a liquid? How could water be heated above its ordinary boiling point? And how is the boiling point affected as we ascend a mountain?

6. If you dip your hand into lukewarm water, and then expose it to the air, the hand feels cold. If you make the same experiment with ether, the hand feels much colder on exposure. Explain these facts.

7. Explain why, in order to cook food by boiling, at the top of a high mountain you must employ a different method from that used at the sea-level.

8. I once went into a room, the doors and windows of which had been kept shut for some time, and the temperature of which was 80° F. I took some water, also at 80° F., and sprinkled it over the floor, and the temperature at once fell several degrees. How do you explain this?

9. A glass water-bottle is covered with felt; explain why the contained liquid may be cooled by moistening the felt and placing the bottle in a dry room.

10. How does a change of pressure affect the temperature of the freezing point of water?

11. How is the temperature of boiling water, and of the steam produced, affected by (i) a change of atmospheric pressure, (ii) a salt dissolved in the water.

12. What is meant by a unit of heat? What is meant by the statement that the latent heat of fusion of ice is 80?

13. You are given a piece of wax, the exact composition of which you do not know. Describe exactly the method of

determining its melting point, which is proved by a preliminary experiment to be below 100° C.

14. What is the effect of dissolved salt on the freezing point of water?

15. Explain why water-pipes burst in cold weather. On freezing water in a glass tube the tube sometimes breaks. Why is this?

16. The images on gold and silver coins are stamped; good castings cannot be taken. Why? Could you obtain a sharp casting of ice?

17. How do you explain the fact that two pieces of ice pressed together unite to form one piece when the pressure is relieved?

18. Explain what is meant by *regelation*, and describe an experiment to illustrate the phenomenon.

CHAPTER XIV

VAPOUR—VAPOUR PRESSURE

Water vapour always present in the air.—The atmosphere always contains a quantity of invisible water vapour, owing to the fact that water exposed to the air evaporates at all ordinary temperatures. This vapour is supplied originally by evaporation from the surfaces of the seas, lakes, and rivers. The amount present in the air is constantly varying; its quantity depends upon the temperature—the higher the temperature the greater the quantity of vapour the air can contain. Further, it can be proved experimentally that *a given volume of air at a certain temperature can only contain a certain fixed quantity of water vapour;* for example, 1 cubic metre of air at 10° C. cannot hold more than 9·3 gm. of water vapour, 1 cubic metre of air at 15° C. cannot contain more than 12·7 gm., whilst 1 cubic metre of air at 20° C. can contain 17·1 gm. of water vapour, and no more.

When the air, according to its temperature, contains the greatest amount of invisible water vapour it can hold, it is said to be **saturated.**

Thus evaporation goes on so long as the air just above the surface of water is unsaturated. Therefore, during winter the streets remain wet for a longer time than in summer, since in the winter the air sooner becomes saturated. Also, evaporation takes place more rapidly when the air around is agitated; for this brings fresh unsaturated air to take up the water vapour. Thus a wind assists in the drying of wet clothes.

When the air is nearly or quite saturated with water vapour it feels damp. The air is said to be dry when the amount of water vapour present is very much less than the quantity required to saturate it; thus *in dry air evaporation proceeds rapidly.* There is usually more water vapour present in the air in summer than in winter, because then the air is at a higher temperature and is thus capable of holding more, though it may feel dry.

Expt. 125.—Place a piece of caustic soda in an evaporating dish. Notice that it soon becomes moist on exposure to the air, because of its absorption of water vapour from the atmosphere. It is for this reason said to be **deliquescent.**

Expt. 126.—Weigh a U-tube containing pieces of calcium chloride or pumice-stone soaked in sulphuric acid. Then, by means of a bellows, pass a current of air through the tube for some time. Weigh again and notice that the weight has increased on account of the water vapour absorbed from the air.

Condensation of atmospheric water vapour.—Condensation is the reverse process of evaporation; it is the return of vapour to the liquid state, and is brought about either by a fall in temperature or an increase of pressure. The condensation of the water vapour of the air is nearly always caused by a fall in temperature.

If the temperature of the air when containing an amount of water vapour insufficient to saturate it be lowered, the air will at length become saturated, and the temperature at which this occurs is called the **dew-point.**

The temperature at which the air is saturated with the water vapour it contains is called its dew-point.

If the temperature of the air be further lowered, the water vapour present is more than sufficient to saturate it and the *excess* of vapour condenses, *the air remaining saturated.* This condensation of atmospheric water vapour results in the formation of dew, hoar-frost, fog, mist, cloud, rain, snow, or hail, according to circumstances.

Dew.—After sunset the earth loses the heat which it has absorbed during the day and cools more rapidly than the atmosphere. As the earth cools the air immediately above it also cools. Consequently a temperature is reached at which the lower layer of air can no longer hold its water vapour; condensation follows, and the excess vapour is deposited as dew.

The greatest quantity of dew is deposited on those surfaces which lose heat most rapidly and which therefore very quickly fall to a low temperature. For this reason very much more dew is deposited on blades of grass and leaves of trees than upon gravel-paths or stones. Anything which prevents the earth from freely losing its heat, such as clouds or trees, prevents condensation. Hence on a cloudy night less dew is formed than on a clear night, and less on the grass under a spreading tree than on the grass in the open.

If the temperature of the surface of the ground, or of other surfaces falls below 0° C., and if there is not much water vapour in the air so that the dew-point is also below 0° C., the water vapour freezes as it is condensed, and is deposited as **hoar-frost.**

Cloud—Mist—Fog.—When a mass of air containing a large quantity of water vapour is cooled to condensation point, some of the invisible vapour condenses into minute particles of water and becomes visible. If this takes place near the surface of the earth it is called **mist** or **fog**; if in the higher regions of the atmosphere it is called **cloud.** The "steam" that issues from the chimney of a locomotive or from the spout of a kettle is not water vapour, but a true mist or cloud.

Rain.—When the exceedingly small drops of water forming a cloud unite to form larger drops, they fall more rapidly, and, if the lower parts of the atmosphere are saturated, they get larger and larger and fall as rain. Sometimes, however, the tiny particles of water as they slowly fall meet unsaturated layers of air and undergo complete evaporation.

Hail—Snow.—If the raindrops form and then fall through colder layers of the atmosphere and are frozen, they fall as hail. If the condensation of water vapour takes place at a temperature below 0° C. the vapour freezes as it is condensed and exceedingly minute ice crystals are formed which grow as they fall. The union of such small ice crystals produces **snow-flakes.**

Any cold body lowers the temperature of the air in immediate contact with it; and, when the temperature is sufficiently lowered, a thin film of moisture is deposited upon the cold body.

Expt. 127.—Thoroughly dry the outside of a flask and then fill the flask with ice-cold water. Notice that moisture quickly collects on the surface of the cold glass.

A person wearing spectacles finds, on entering from the cold out of doors into a very warm room, that his spectacles are bedewed because the warm air—laden, but not saturated, with moisture from human breath—which is in immediate contact with the cold spectacles is chilled, and at this lower temperature cannot retain all its vapour. Some is therefore condensed and deposited as dew on the spectacles.

The presence of the water often found trickling down the inside surfaces of the windows of a warm room on a very cold day is explained in a similar manner.

Experiments to illustrate properties of vapours.

Expt. 128.—Fit up a simple mercurial barometer according to the instructions given in Expt. 92, page 53, and mark on the tube, by means of a piece of gummed paper, the level of the mercury surface. Then introduce a drop of water into the tube. This may be accomplished by means of a pipette, the lower end of which has been softened in a blowpipe flame and then bent round. Suck some water into the pipette, and then place the short curved portion into the end of the barometer tube; blow gently into the top end of the pipette so as to cause a drop of water to rise through the mercury column into the Torricellian vacuum. Notice that the water evaporates rapidly and disappears. Introduce another drop of water—it evaporates and the vapour formed exerts a downward pressure on the mercury column, causing it to fall slightly. Note the height of the column.

Introduce more water, when the column falls still further, showing that the space above the mercury in the tube is still unsaturated with water vapour. Continue in this way until some water remains above the mercury; the space is then saturated with water vapour *at the temperature existing at the time of the experiment*, and the mercury surface ceases to fall, showing that the vapour is exerting its **maximum pressure at that particular temperature.** Carefully measure, in mm., the height of the mercury column when the space is just saturated, and subtract the result from the original height of the barometer to obtain the maximum pressure of water vapour at that particular temperature.

If the size of the saturated space be reduced the pressure of the water vapour does not alter, but some of the vapour condenses to the liquid form; if the volume of the space be increased more of the liquid at the top of the mercury evaporates, until, if there is sufficient liquid, the space again becomes saturated and the maximum pressure reaches the same value as before.

Expt. 129.—Repeat the previous experiment with alcohol instead of water. Notice that the maximum pressure exerted by the alcohol vapour is greater than that of the water vapour—the mercury column falls further when the space is saturated with alcohol vapour.

Expt. 130.—Repeat Expt. 128, using ether in place of water. Notice that when the space is saturated with ether vapour a greater fall of the mercury column is obtained than when saturated with alcohol vapour.

The fall is thus greatest for ether vapour and least for water vapour; *i.e.* at a given temperature, the maximum vapour pressure of ether is greater than that of alcohol, and that of alcohol greater than that of water.

If the temperature of the saturated space be raised more vapour is required to keep it saturated; if the temperature be lowered less vapour is needed to produce saturation, some of the vapour condenses, and the maximum vapour pressure falls in value.

Expt. 131.—Fit up the apparatus indicated in Fig. 53, using one barometer tube instead of two. Note the height of the mercury column and also the temperature of the air. Introduce liquid into the vacuum until the space is saturated and observe the depression produced. Now connect up the outer glass tube or jacket A to a water trough so that the barometer tube may be surrounded with water at different temperatures. Heat the water trough, and when the temperature reaches 20° C. allow the water to run through the jacket tube. Note the exact temperature of the water jacket by means of a thermometer attached to the barometer tube. Observe that at this higher temperature more liquid will evaporate into the space, and also that a further depression is produced.

Repeat the above at higher temperatures, and it will be found that *the higher the temperature the more vapour is needed to saturate the space.*

Finally, run cold water through the jacket tube and notice that some of the vapour condenses and the mercury column rises.

In the preceding experiments the liquid was allowed to evaporate in a vacuum.

Expt. 132.—Fit up a simple mercurial barometer and introduce a small quantity of air by blowing gently through the pipette. Then force a drop of water to the top—it evaporates and the column is depressed. Introduce more water and notice that finally a limit is reached at which evaporation ceases and a thin film of water remains on the surface of the mercury. The space containing air is then saturated with water vapour at the existing temperature.

Carefully note that in the presence of the air evaporation proceeds more slowly than in the vacuum.

EXAMPLES XII

1. What is meant by the statement that the dew-point is 8° C.? How is dew formed, and why is it more copious on some substances than on others?

2. When is a space said to be saturated with vapour; and what is the effect of (*a*) a fall, (*b*) a rise in the temperature of a space which is so saturated?

3. What circumstances determine whether a damp towel exposed to the air shall dry or shall remain damp? Why does a damp cloth exposed to draught become very cold?

4. Explain the formation of cloud, snow, and hoar-frost.

5. What becomes of the water of puddles when they "dry up"? What conditions of the air are most favourable to the process?

6. Explain the fact that no dew is found upon the ground after a cloudy or a windy night.

7. A glass bottle and a bottle of porous earthenware are

both filled with water and exposed to the air side by side. Usually the water in the earthenware bottle becomes decidedly colder than that in the glass. Why is this? If there is little or no difference of temperature, what conclusion may we draw as to the state of the atmosphere? and why?

PART III

CHEMISTRY

PART III

CHEMISTRY

CHAPTER I

THE IDEA OF CHEMICAL CHANGE

BEFORE we turn to the chemistry section of our syllabus it is essential to realise what is meant by a chemical change.

It is obvious that most substances and objects we meet with are continually changing, although in different ways. For example, coal burns away leaving a whitish ash. The same coal, left exposed to the air for a great period of time, would disappear.

Ice on being warmed melts to water; water when frozen takes on the solid nature of ice. Candles apparently vanish on burning. Milk exposed to the air turns sour, and we ourselves are subject to these inexorable variations.

Are these various kinds of change really the same, or can we distinguish between them?

To answer this question we must make **experiments**, that is to say, allow different changes to take place, and carefully observe what happens.

Expt. 1.—Heat a small fragment of **sulphur** in an iron spoon or on the point of a knife-blade. The sulphur burns with a pale blue flame and gives off a suffocating fume. No residue remains.

Expt. 2.—Moisten some **iron filings** and allow them to stand exposed to the air for some time. They gradually become rusty. This rust can be scraped off, and certainly bears no resemblance to the iron. The iron seems to change into quite a different substance.

Expt. 3.—Heat a little **sugar** in an iron spoon. The sugar melts, catches fire, and leaves behind a black, crisp material. On further heating this also glows and disappears.

Many similar experiments might be performed, but the above examples show that *some changes cause a complete alteration;* the original body either utterly disappears or leaves behind quite a different substance.

Such a change is called a **Chemical change.**

It is the object of Chemistry to examine such changes and to find out what is formed after a chemical change.

But are chemical changes the only kind we meet with? When ice is formed from water by freezing, have we a chemical change to deal with? Ice, as we know, can easily be changed back to water by simply reversing the process. Could we in Expt. 3, for instance, turn the black ash into sugar by reversing the operation, *i.e.* by cooling? Evidently not. This change of water to ice, then, seems of a different nature from the preceding examples.

Let us see if we can find other instances.

Expt. 4.—Rub a steel knitting-needle or knife-blade with a magnet. It will now attract iron filings, a property which it did not possess before. It has been changed to a magnet. But it is still the same piece of steel. If left to itself for some time, or heated and allowed to cool again, it will lose this magnetic property. Evidently this is not a chemical change.

Expt. 5.—Heat a piece of **platinum** foil in a bunsen flame. It has now new properties, which it did not previously possess; *e.g.* it glows brightly, giving out light and heat. But on cooling it returns to its original form and appearance. This change, then, can be reversed, and throughout the change the platinum has remained platinum. Heat a few cms. of **magnesium** ribbon, and observe what changes take place.

Such changes as these, in which no new bodies are formed, but only new properties added, and which are easily reversible, giving back the unaltered substance, are called Physical changes.

It is usually difficult, and in some cases impossible, to reverse a chemical change.

As further instances of these changes the two following experiments may be quoted.

Expt. 6.—Place some lumps of freshly ignited quicklime in a dish. Pour on the lime a little cold water. Presently the lime breaks up into powder, steam is evolved, and quite a vigorous action goes on. Slaked lime is formed.

Expt. 7.—Heat some **sal-ammoniac**—a white crystalline powder—in a clay dish. The sal-ammoniac does not melt, but, some distance above the dish, white fumes are seen. Hold a cold surface, such as a knife-blade, in the fumes.

A white solid is seen to be deposited. Bring the flame of your bunsen burner in contact with the fumes. They at once disappear, but are formed again when the flame is removed. Heat the white deposit on your knife-blade before again it disappears and form similar fumes. Taste the deposit and compare it with that of the original substance. A change has ensued which obviously has not resulted in the formation of any new substance.

Expt. 8.—Take a porcelain crucible provided with a lid. Place in the crucible some small pieces of **magnesium** ribbon.

Accurately counterpoise the crucible and contents by placing them on the balance pan and adding shot to the other pan till the balance is in equilibrium.

Place the crucible on a pipe-clay triangle, supported on a tripod stand (Fig. 1), and heat with a bunsen flame.

Slightly raise the lid from time to time with the crucible tongs, to allow access of air. The magnesium burns, forming a white ash, which would escape were it not for the lid. When no more burning takes place, remove the lid, and heat strongly for a time. Replace the lid and allow the crucible to cool till it can be easily handled, and again place it on the balance pan. *A considerable increase of weight will be noticed.*

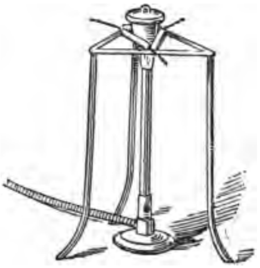


FIG. 1

Here, then, we see that a chemical change has gone on and there has been an alteration in weight, the residue being heavier than the original substance.

Summing up these introductory words we may state, **after a Chemical change has taken place :—**

1. There is a complete alteration in properties and appearance, a new body being formed.
2. Some heat effects are noticed. Frequently heat is evolved.

3. A change in weight is often observed.
4. There is often much difficulty in reversing the operations and obtaining the original substance.

Whereas, **after a Physical change**:—

1. No new substance is formed.
2. No change in weight has been observed.
3. It is possible to reverse the operations and obtain the original body unchanged.

CHAPTER II

SOME COMMON SUBSTANCES AND THE EFFECT OF HEAT ON THEM

THE different behaviour of substances when heated is often characteristic—sometimes a purely physical change ensues, in which no new substance is formed but fresh properties are for the time being conferred, as for instance the turning of ice into its liquid form—water. Or on the other hand a chemical change immediately sets in, as when gunpowder is ignited; again, it is possible for both varieties of change to operate when we first melt the wax of a candle, then turn the melted wax into a gas, and finally set fire to this gas—all three changes happening when we light a candle.

We will now consider the effects of heat on a variety of common materials.

Wood.—Hold a splinter of wood in a bunsen flame. After a short time it catches fire and will, if held in the air, presently burn away, evolving much heat and light—evidently a chemical change is going on. During the burning blow out the flame and notice that an irritating smoke or fume arises from the wood, the top of which still smoulders.

Now all this time the wood has been burning freely exposed to the air. What will take place if we heat the wood in some vessel from which the air has been excluded?

Expt. 9.—To answer the question place a few chips of wood on an iron tray and cover them with sand. Heat the tray. As the sand tray gets hot a smoke arises through the sand, but no flame. After some minutes allow the tray to cool. When cold remove the sand and examine the wood. It is now black charcoal, and, as is well known, wood can be readily converted into charcoal by covering heaps of it with turf and brushwood to almost exclude air and then allowing part of the heap to burn and evolve sufficient heat to carbonise the rest.

A better way of heating substances away from any influence of the air is by the use of a bulb-tube.

Expt. 10.—To make a bulb-tube. -

Take a piece of glass tube about 6 inches long, and round the rough edges by holding it in a bunsen flame *for a few seconds*. Hold it with both hands in the bunsen flame so that it is heated in the middle. Keep the tube slowly moving in the flame by rotating it.

Never heat glass rapidly in one place, always **keep it moving in the flame** till thoroughly hot.

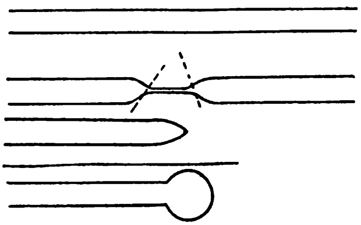


FIG. 2

Presently the tube softens in the middle. Take it from the flame and pull out each end horizontally. The softened part is drawn out into a capillary. Break this off about 5 cm. from the end, and then heat the tube just where the capillary joins on. Pull off the capillary when the glass is soft, leaving a closed tube (Fig. 2). Hold the closed end in the flame, constantly rotating it, till it is quite soft. Remove the tube from the flame, hold it downwards, blow steadily into it, still rotating it.

The softened end is thus expanded into a bulb.

Having thus made a bulb, place in it a scrap or two of wood. Warm gently in the bunsen flame. Notice that fumes are given off which partly condense or assume the liquid form when they come in contact with the colder sides of the tube. The fumes which escape have a characteristic odour, and if a piece of blue litmus paper be held at the mouth of the tube it is soon turned red. Litmus is a vegetable colouring matter, normally blue, but turned red by acids. So that the heating of wood away from the air causes the evolution of acid fumes, and a black residue of charcoal remains behind. Such a process, when a body is entirely broken up by heat into gases and vapours, is known as destructive distillation.

Thus whether the wood is heated in the presence or absence of air a very marked chemical action proceeds, but the action varies very considerably with the condition.

Paper.—Seeing that paper is closely allied to wood, and in fact most of the paper we now use is made from wood-pulp, we should expect that very similar action would proceed whether we heated wood or paper. Why does a piece of paper burn more rapidly than a wooden stick? Evidently because there is more chance of the air reaching every fibre of it, for we have just learned that without air wood will not burn, but destructively distils and leaves charcoal.

Try then whether a scrap of paper heated in a bulb-tube leaves charcoal or not. You will have noticed in the cases of both wood and paper that when burnt in air a whitish or greyish ash is left. This is the mineral matter which the tree or plant from which wood or paper was derived had absorbed from the soil on which it grew. In the days before the introduction of soap these wood ashes or pot ashes were of great use in washing, and from this fact we derive the name potash and potassium, the metal contained in potashes. Place some of the wood ashes you have made on reddened litmus paper. Notice that they turn red litmus blue. The Arabic name of wood ashes was alkali, and now we term alkaline all substances which turn red litmus blue, just as acid (or sharp-tasting) substances turn blue litmus red. Litmus is thus an **indicator** for acids and alkalies.

Coal is the relic of the marshy forests which existed countless years ago. The microscope teaches its vegetable origin, so that once more we should expect that the heating of coal would not be very unlike that of wood and paper. On a large scale coal is used for heating purposes. In the air it burns freely, some varieties burning with a smoky flame and sending out jets of gas, others with a clear red heat like coke. In each case an ash or cinder remains after the chemical change is over.

But if the heating is carried out in closed iron boxes or retorts away from free access of air, then we have a large amount of an inflammable gas—coal gas—given off, together with tar and a strong-smelling, watery ammoniacal liquor. This process is carried out in the gas-works on a colossal scale.

Expt. 11.—Place some fragments of coal in a bulb-tube and heat strongly. Notice the smoke which issues and by applying a light see that this smoke or fume is inflammable, being indeed impure coal gas. All the while tarry matter has been forming on the side of the tube, and with it drops of ammoniacal liquor.

Here again we have a process of destructive distillation marked by very profound chemical changes—heat and light have been evolved during the burning of the volatile products of the distillation.

Starch is one of the first products of plant life. During the day green plants are able to assimilate carbonic acid gas from the air and water from the soil, and by some wonderful chemical agency to transmute these simple inorganic substances into highly complex organic starch.

A slice of potato scraped under water soon causes a milkyness due to the escape of the tiny starch granules.

On being heated freely exposed to air starch becomes brown and then chars without burning. As would be expected from its vegetable origin, the fume evolved is inflammable, and on further heating the starch leaves a crisp black ash of charcoal which slowly smoulders away.

Again, heat starch in a bulb-tube and observe the almost identical destructive distillation that goes on.

Sugar is produced in the vegetable organism subsequent to the production of starch. A piece of starch placed on the tongue soon tastes sweet. The saliva contains a ferment which is able to convert starch into sugar.

Perhaps the most striking difference between sugar and the vegetable products above mentioned, viz. coal, wood, paper, and starch, is its crystalline structure. That is to say, if a particle of sugar is examined it is found to be bounded by regular facets. These facets are square, and the particle is a cube; moreover, any particle of sugar is also cubical. A crystal, then, is a regular geometrically shaped substance, and is usually formed by the cooling or the gradual evaporation of a solution. Cane sugar has been crystallised from the sap of the sugar cane; beet sugar from the purified juice of beet. If sugar is gently heated in a clay dish a very striking series of changes is to be noticed. The sugar browns and then melts to a sticky mass of caramel. Further heating causes a bubbling due to an evolution of white gas, and this gas soon catches fire and burns with a clear luminous flame, leaving a black crisp mass of charcoal. This gradually smoulders away. By heating the sugar in a bulb-tube we again have a destructive distillation and a very pure variety of charcoal remains.

Hitherto we have studied in this chapter the heating of vegetable matter. We now turn to an animal product.

Feathers when heated shrivel up and evolve a most unpleasant odour. They do not burn so readily as wood, for instance.

If red litmus paper is held in the fume it immediately turns blue.

Expt. 12.—Place some feathers in a test-tube and gently heat them. Notice the objectionable smell emitted and the condensation of tarry matter on the sides of the tubes. Notice also the alkaline nature of the gases evolved. This destructive distillation also results in the formation of a variety of charcoal—animal charcoal—which is, however, much less pure than wood charcoal. Very similar facts are noticed when leather, skin, or hair are heated either in the open air or in bulb-tubes. Unpleasant alkaline gases are evolved, tarry liquids distil, and an impure carbonaceous mass remains.

CHAPTER III

THE EFFECT OF HEAT ON SOME METALS AND SALTS

In the last chapter we considered the differences which appeared when various organic bodies were heated. We now pass to certain inorganic substances, viz. salt, soda, chalk, copper, iron, and lead.

Three of these six materials fall under the heading of metals—copper, iron, and lead. Metals are usually dense, opaque solids, possessing a distinct lustre or sheen. They are frequently malleable or able to be beaten out into thin sheets, ductile or able to be drawn into wire, and tenacious or possessing a strong tendency to resist rupture. Copper possesses these three properties to a very marked extent, whereas lead, although malleable, is not particularly tenacious.

Expt. 18.—Take a few scraps of lead—notice that the metal is so soft that a little pressure will flatten it. It can be cut easily with a knife and will mark paper. It is malleable and fairly ductile. Heat the metal on a clay dish. It soon softens and melts to a shining silvery mass. Careful watching will show you that after a time a play of colours appears on the shining surface. This goes on until the bright metal is quite covered with a brownish scum, which on cooling appears yellow. Continue the heating till all the lead is converted into this dross or earthy yellow substance. It is known as litharge.

Now, why should lead on being melted and heated turn into litharge?

From what has gone before we have found that the air makes a very considerable difference during heating, so it appears likely that if we heated the lead away from the air this formation of litharge would be hindered, just as wood is unable to burn when it is covered with sand.

Try similar experiments with iron and copper. Iron is a grey-white metal, which, unlike lead, rapidly consumes away when exposed to air and moisture. We say that the iron **rusts.** But it has been found that if iron is covered with a layer of zinc or galvanised that this rusting goes on much less rapidly—so that if protected from air iron does not rust. Similarly iron can be kept bright if placed under water which has been recently boiled to free it from dissolved gases.

Iron possesses the qualities of durability, malleability, and tenacity to a very marked extent, and these properties make it the useful metal it is. It is rather hard to melt, but when white hot can be **welded,** that is united firmly, to another similarly heated mass of iron.

Notice that when you heat scraps of iron or iron filings in a clay dish a black layer is formed. The same black substance appears when a blacksmith hammers out his red-hot iron. The so-called blacksmith's scales are formed, just as litharge, by the action of the air on the hot iron. What would happen if iron were heated away from the air? Would you notice any such black deposit?

Copper is a ruddy golden metal possessing to the highest degree the properties of durability, tenacity, and malleability.

It does not easily melt and is fairly heavy. Very little corrosion goes on when copper is exposed to the air, consequently it is a suitable metal for coinage, whilst being a good conductor of heat and electricity it serves for boiler tubes and telegraph wires.

Expt. 14.—Place a scrap of copper turnings on a clay dish and gently heat. Watch carefully for any play of colour such as you obtained in the case of lead. The play of colour will be due to the formation of a thin layer of some such substance as litharge, or blacksmith's scales.

After continued heating the copper will be found encrusted with a black material which can be scraped off, leaving still unaltered copper beneath.

Now heat a further amount of copper in a bulb-tube. Notice that as the air is practically excluded no such formation of black crusts takes place, the metal merely becoming red hot.

Thus the general lesson we learn from these three metals is that on exposure to air more or less tarnishing ensues, resulting in the formation of an earthy crust or dross, and that in the case of iron this rusting is very pronounced. Lead also usually has some such dross on its surface, but this does not always result in the complete corrosion of the metal. Copper is the most permanent of the three. Gold and silver are remarkably permanent, thus earning their names of noble metals.

We shall see later what an important stride it was in the history of chemistry when this question of the tarnishing of metals was first really understood.

We now pass to the behaviour of three very common substances, viz. salt, soda, and chalk.

Salt is one of the commonest substances in use, and occurs universally in nature, and particularly in solution in the sea.

If you can obtain any rock salt examine it and notice its crystalline nature. Perhaps you will be able to observe the cubical shape of its crystals. It is characterised by a saline taste, and is the best-known member of the great chemical family of salts—substances of a somewhat saline taste—soluble in water, and usually having no effect either on red or blue litmus.

If salt be strongly heated it melts to a clear liquid which solidifies on cooling. Such fused common salt is much used in the laboratory.

Soda, or washing soda, is another salt of very common occurrence, being manufactured in immense amount from common salt.

Obtain a fresh specimen of washing soda. Notice its clear crystalline structure and that it feels peculiarly soapy. Place some in water and notice how soluble it is. This solution is a good solvent for grease, hence its use in cleansing.

If you take a crystal which has been standing exposed to air for some time you will observe that it has turned on the surface into a white powder which is no longer crystalline. A substance which does not possess crystalline form is said to be **amorphous.** Thus on exposure to air washing soda becomes amorphous.

Heat washing soda on a clay dish. Bubbles are seen. The soda melts and steam is evolved. A white amorphous powder remains. So that on heating soda loses water. This loss of water causes loss of crystalline form, and you will find that many other crystalline substances evolve water on being heated. Such water is called **water of crystallisation.** After loss of this water an anhydrous solution remains which is also amorphous.

If you heat soda in a bulb-tube you will see that this water of crystallisation will be condensed on the cold sides of the tube, and a considerable amount of liquid will be deposited.

Chalk, like common salt, is a material of widespread distribution in nature. A great portion of the surface land of England is of the chalk formation and of its chemical alias limestone.

A common form of chalk is known as **whitening**. A purer form is the precipitated chalk of the druggist.

Thus seen, it is a white, very fine powder, insoluble in water, and without any effect on either red or blue litmus.

Expt. 15.—Take a little chalk in a test-tube and add a few drops of any dilute acid. There is a vigorous effervescence of a gas. It will be found that a lighted splint or taper will be extinguished when held in the gas. The gas thus made is known as **carbonic acid** gas.

Expt. 16.—Find if the same gas is set free when acids are added to limestone or washing soda.

Heat a little chalk on a pen-nib or knife-blade. Notice that when strongly heated the mass becomes luminous or incandescent. When cool place this white residue on moist red litmus. A blue stain is formed, hence the effect of heat on chalk is to form an alkaline substance. Test whether this new substance effervesces with acid. Also try a little with a drop of water. Notice the steam evolved. The substance formed when chalk is strongly heated is known as

Quicklime, which is an alkali and absorbs water eagerly, giving off heat and forming slaked lime.

CHAPTER IV

THE ATMOSPHERE

ALTHOUGH the air has been recognised from the earliest ages as necessary for life, it is only within comparatively recent times that its true importance has been known.

What happens when a candle burns away in air? Is it really lost? Has the air around played any part in the burning?

Again, why does lead turn into letharge, copper into a black ash, and iron into red rust, while away from air none of these events take place?

About the middle of the seventeenth century it was noticed by **Hooke**, a celebrated chemist at Oxford, that wood could not be converted into charcoal in the open air. The air must be excluded, or the wood simply burns away. If, however, the wood be covered with sand, or, on the large scale, with turf, and then be heated, charcoal will be formed. Again, about the same time, it was shown by **Boyle**, an early experimenter on gases, that substances such as sulphur, which burn quite easily in the open air, will not burn when heated in a vacuum.

So that to some men of this period it was certain that the air took a very important part in cases of burning, and indeed that it was *an essential factor.*

Now this question of burning, and more especially the burning or **calcination** of metals, attracted very considerably the attention of the chemists of this time, and an explanation was offered, which at first sight appears very reasonable.

This explanation asserted that combustible bodies, *i.e.* those which burned when heated in air, contained a principle or spirit of fire, which was termed **phlogiston**, from a Greek word meaning to burn.

This phlogiston was assumed to escape during combustion, and its escape was usually attended by flames and emission of light and heat.

The residual body in the case of metals was called the **calx**, or ash.

Thus when we heated magnesium ribbon (p. 99), this phlogistic theory or explanation would say that when the magnesium was burning the phlogiston was escaping, leaving at last the white calx or ash in the crucible. Oil, wax, coal and wood were said to contain much phlogiston, because they burned so easily. The calces of metals, on the other hand, contained no phlogiston. How, then, could a metal be obtained again from its calx? Obviously by *adding phlogiston to the calx.*

Exp. 17.—To make a metal from its calx.

Mix some **litharge** (the yellow calx of lead) with powdered charcoal and heat the mixture strongly in a crucible. Beads of lead are produced. Now, by the phlogistic theory, charcoal contains much phlogiston, and so this phlogiston has added itself to the lead calx forming metallic lead.

Notice that this theory quite ignores the possible effect of the air, and also is not based on any weighings and measurements. In fact, the early chemists entirely disregarded accurate estimations in their inquiries, and it was not till modern times that the balance began to take such an important place in the sciences.

Now let us go into this question of burning more fully, and see what conclusions we can come to after trying several experiments.

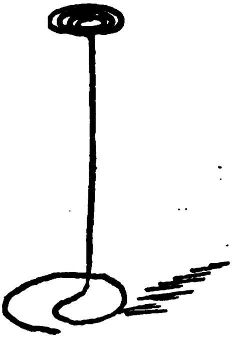


FIG. 3

Exp. 18.—To burn phosphorus in a closed space of air.

Take a piece of stout copper wire, 20 cm. long, and twist one end into a circle so that the wire will stand up. Make a little spiral or platform at the other end by twisting the wire round several times (Fig. 3). A wire stand is thus formed on which the various substances we are going to heat can be supported.

Place this stand in a trough or sink containing water to the depth of a few cms.

Place a fragment of **phosphorus**—a *yellow, waxy, very inflammable solid*[1]—on the top of the stand, and set light to it.

Rapidly and carefully bring a dry jar over the stand, and allow it to rest on the bottom of the sink. We then have the phosphorus burning in a closed space of air (Fig. 4).

Notice (1) the dense white fumes evolved by the burning phosphorus.

(2) That the phosphorus presently goes out, but that there is usually some unburnt phosphorus left behind so that the residual air will not allow burning to proceed in it.

(3) That water rises, filling about one-fifth of the jar.

(4) That presently the white fumes dissolve in the water.

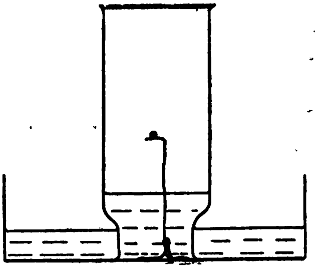


FIG. 4

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[1] This substance must be kept under water because of its inflammability. It should be cut under water, rapidly dried, and placed on the stand. Avoid handling it overmuch.

Red phosphorus may be used. This is not poisonous and not very inflammable.

Incline the jar, keeping its mouth under water, and withdraw the wire stand.

Place your hand under the mouth of the jar and take it out of the sink, carefully avoiding any loss of the water inside.

Pour the water in the jar into a measuring cylinder. Find the volume it occupies, taking care that when reading the graduations your eye is on a level with the surface of the water in the jar (A, Fig. 5).

Another point to be observed in reading the volume of a liquid in this way is to read always from the bottom of the curved surface formed, which is called the **meniscus**. In the case of mercury the meniscus is convex, and then you must read from the top of the curve.

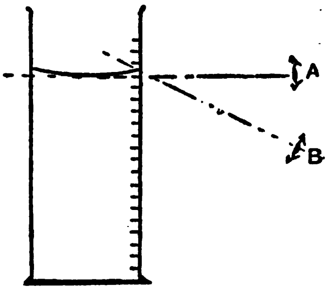

FIG. 5
A, correct position.
B, incorrect position.

Exp. 19.—Repeat Exp. 18, using instead of phosphorus a small piece of **taper** fixed to the top of the wire stand. Notice a mist appearing on the surface of the jar, while the taper is burning, and that again there is a rise of water about one-fifth up the jar.

Exp. 20.—Do a similar experiment, using **sulphur.**

Now measure the volume of the jar, by filling it with water and pouring the water into a graduated measuring cylinder, taking the precautions in reading mentioned above. Find the percentage of the volume of air in the jar which has disappeared during the burning of the phosphorus, taper, and sulphur.

The following average percentages were obtained by classes of students doing the above experiments:—

	Exp. 11	Exp. 12	Exp. 13
Class A	23	23	20
„ B	21	23	22
„ C	23	23	21

It seems, then, that when these three substances burn in a closed space of air, **in each case about one-fifth of the air vanishes**, and is replaced by water.

What has become of the air? Is it possible that the burning phosphorus, for instance, has united with it to form the dense white fumes we noticed?

If this were the case, would there be any difference between the weight of the fumes and the weight of the original phosphorus? Evidently there would be a difference, due to the weight of the air which has united with the phosphorus.[1]

Let us test the supposition by experiment.

We have to arrange our apparatus so that the phosphorus we are going to burn, and also the white fumes formed, can be weighed.

Exp. 21.—To find if the fumes given off by burning phosphorus are heavier than the phosphorus.

Take a piece of stout glass tubing about 15 cm. long and 2 cm. in diameter. Round off the edge in a bunsen flame.

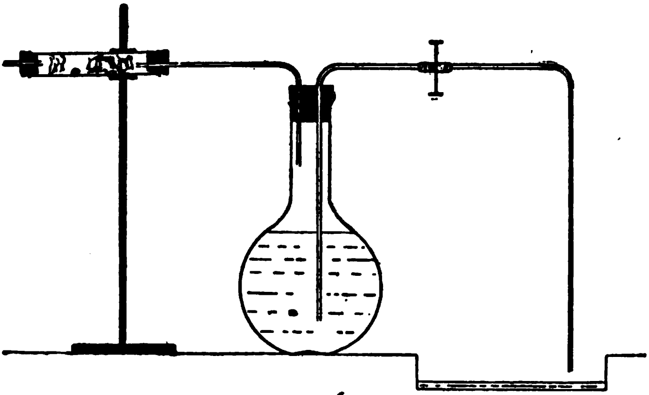

FIG. 6

Provide it with corks which have been previously softened. Bore a hole in each cork and fit a straight piece of glass tubing about 6 cm. long into each hole (Fig. 6).

Loosely pack half the tube with **asbestos wool**. This serves as a filter to retain the white fumes.

Now take a large flask containing nearly a litre of water. Fit it with a softened cork which is bored with two holes. Through one hole put a short tube, bent at right angles (Appendix), and through the other hole put a long right-angled bent tube passing nearly to the bottom of the flask. Connect the projecting end of this tube with a similar longer right-angled bend by indiarubber tubing, which carries a spring clip.

A siphon is thus formed. Fill it with water by blowing through the short right-angled bend and then close the clip.

Now place in the large tube a small piece of **phosphorus**, introduce a little more asbestos wool, put in the cork and counterpoise the whole on the balance pan. Support it on a retort stand, between two rings, and connect up with the flask by rubber tubing.

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[1] According to the Phlogistic Theory the fumes should weigh less than the phosphorus.

Open the clip and allow water to run out of the flask. Air enters the tube to take the place of the water in the flask. Gently warm the phosphorus with a bunsen flame. It catches fire and the white fumes formed are nearly all caught in the asbestos wool filter.

Disconnect the apparatus when the burning ceases, and allow the tube to cool. Place it again on the balance pan, when a considerable gain of weight will be observed.

It seems reasonable to suppose, then, in this experiment, that the increase of weight after burning is due to the fact that **air is taken up by the phosphorus.** Moreover, it is probable that it is the one-fifth of the air which we found to disappear in Exp. 18. It appears as though this would explain the corresponding loss of air when the taper and sulphur were burned. Many experiments, however, are necessary before we can accept this supposition as a correct one.

Exp. 8 gives our explanation some assistance, for it shows that the calx or ash of magnesium is heavier than the original magnesium, as we should now expect it to be. Can you see any serious objection, which Exp. 8 offers, to the Phlogistic Theory?

Let us now see if other metals also form calces which are heavier, as in the case of magnesium.

Exp. 22.—To find if lead increases in weight when calcined.

Counterpoise some scraps of **lead** in a little clay dish (such as was used in exercises, Chap. I). Support the dish on the edge of a tripod stand, and heat with a bunsen flame. The lead melts, and a **play of colours is noticed on its surface.** A film can be scraped off from the melted lead with the tip of a knife blade, but a fresh film forms immediately, and as it thickens the colours appear again. A similar appearance is obtained when oil is placed on water. If this thin film or layer be constantly removed to the side of the dish, all the lead can be converted to an earthy powder, which, as it cools, becomes yellow in colour, and is known as **litharge** (p. 105)

Place the dish on the balance pan again, and see if any increase in weight can be observed.

EXERCISE

Try if zinc or iron filings give similar results when calcined in air.

We noticed in Exp. 2 that moistened iron filings gradually rusted away in air. Does the air take a part in this process as it does in burning? Let us see what happens if iron filings are allowed to rust in a closed space of air.

Exp. 23.—To find what happens when iron rusts in a closed space of air.

Shake up some **iron filings** with water and get them thoroughly wet (it may be necessary to add a little **caustic soda** to cleanse them from grease). Float a crucible on water which is standing a few cms. deep in a trough.

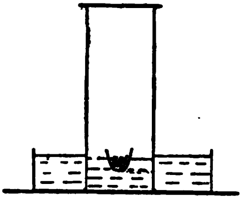

FIG. 7

Place in the crucible as much of the filings as it will carry. Carefully place over it a jar as in Exp. 18 (Fig. 7). Allow the apparatus to stand in a warm place for several days. Notice that the water gradually rises in the jar. When the filings are thoroughly rusty no further rise takes place.

Withdraw the jar as before. Place a lighted taper in the residual gas in the jar. The gas left will not allow the taper to burn.

Measure the amount of water in the jar, and also find the volume of the jar. Find the percentage of air which has disappeared. *Is it the same as was found for phosphorus?* If so, it appears that **rusting of iron is a similar process to burning**, only, of course, much more gradual.

EXERCISE

Devise an experiment to find out if the rust of iron weighs more than the metal from which it is formed.

If a dry stick of phosphorus be examined, a kind of white smoke is seen to arise from it. This suggests the idea that the phosphorus is *smouldering or "rusting"* away while exposed to the air.

Perhaps if we allowed it to smoulder or "rust" in a jar of air standing over water we might obtain a result similar to that obtained when iron rusts.

Exp. 24.—To find what happens when phosphorus smoulders in a closed space of air.

Take the same apparatus as in Exp. 18, but tie a small stick of **phosphorus** to the wire stand with thin copper wire. Allow it to remain for several days, noticing that at first the phosphorus fumes continually and that water rises gradually in the jar (Fig. 8). When no further rise is noticed, take away the phosphorus and its support and remove the jar as before. Again test the residual air with a lighted taper. Again it is seen that this gas will not allow a taper to burn in it. Find the percentage of air which has been used up. Compare your results

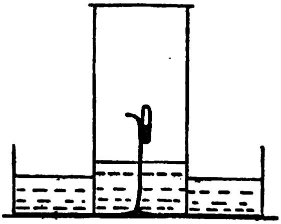

FIG. 8

with those obtained in previous experiments. A similar experiment gave the following numbers:—

Volume of water which rose in the jar = 80 c.c.
Volume of jar = 380 c.c.

$$\therefore \text{Percentage of air used up} = \frac{80 \times 100}{380} = 21.$$

It seems, then, that phosphorus, whether burning vigorously or smouldering slowly, has the same effect on the air. **In each case one-fifth of the air disappears.**

Our supposition that substances which burn in air take up one-fifth of it, and thus form new bodies of greater weight, seems supported by all the experiments we have hitherto made.

Moreover, such changes as rusting and smouldering also bear out this explanation.

Again, in our experiments we find left behind a residual amount of gas, *about four-fifths of the original volume, which seems incapable of allowing burning to take place in it.*

Does it not appear, then, as if air were really a **composite substance,** one part allowing and taking part in burning and the other quite inactive?

If this is so, might we not obtain the active constituent from the various calces or ashes we have produced, or from similar bodies?

It seems certain, then, that one crucial experiment will settle the matter, and that is to form a calx by allowing a body to burn, or even to rust, in air, and then endeavouring to obtain from this calx the active one-fifth of the air.

CHAPTER V

THE SEPARATION OF THE ACTIVE CONSTITUENT OF AIR

WE have seen in the previous chapter that burning bodies "fix," or take up the **active one-fifth** of the air, and that the substance formed, whether fume or calx, is heavier than the original substance.

The problem now before us is to recover this **active air** from some residue left after burning.

Suppose we try the effect of strongly heating some calx.

Red lead is a calx of lead formed when lead is heated for a long time in air, litharge, as we know, being first formed.

How shall we test, if active air is evolved on heating red lead? *Will it not* **support combustion,** *i.e. allow burning to go on in it, very easily?*

Exp. 25.—To find the effect of heating red lead.

Make a bulb-tube (p. 101). Cut a slip of paper which, when folded lengthwise, can be easily slipped down the tube. Place a little **red lead** in the groove thus formed, and introduce the slip into the tube. By gently shaking the red lead will be obtained in the bulb without soiling the sides of the tube.

Heat the bulb in a bunsen flame. Notice that the **red lead darkens in colour,** becoming black eventually.

Hold a glowing splint of wood at the mouth of the tube. It bursts into flame. Evidently a gas is being evolved which supports combustion eagerly. **This gas must be active air.**

On cooling a yellow substance is found in the bulb, and this is obviously litharge.

The experiment we have thus made was first performed, though in a different manner, **in 1774, by Dr. Priestley,** a brilliant chemist, who, in an almost accidental way, discovered not only this but many other gases.

His method in this case was to heat the red lead by condensing on it the sun's heat by a large burning-glass or lens, and he found, as we have done, that a glowing splint of wood inflamed in the gas and rapidly burnt away.

He also found that the red calx of mercury would, on being heated, yield this active air quite easily, while two years before **Scheele,** a Swedish apothecary, obtained the same gas by strongly heating **nitre** or **saltpetre.**

EXERCISE

As in Exp 25, heat nitre, red calx of mercury (oxide of mercury), potassium chlorate, potassium permanganate and potassium bichromate in bulb-tubes, and find out if active air is evolved.

Describe what happens in each case and also describe the various residues left.

Notice, however, that although Priestley and Scheele had each obtained this gas, neither had any clear idea as to how it was related to ordinary atmospheric air; and instead of applying the facts they had discovered to test the Phlogistic Theory, they attempted rather to fit them in with it, not recognising that a theory is made for the facts and not the facts for the theory.

Now there was at that time in Paris an illustrious chemist named **Lavoisier**, who, in great contrast to Priestley, did not experiment in a haphazard fashion, but worked methodically, and made the balance his most important piece of apparatus. Had this been done before, the Phlogistic Theory, as we have seen, could not have survived.

To Lavoisier Priestley communicated his discovery of this active air, and soon Lavoisier, seeing how important this was, made the following experiments.

Lavoisier's Experiment.—Mercury was placed in a flask, **A** (Fig. 9), the neck of which was drawn out and bent up, so that the bend could rest under mercury placed in a trough, **C**, with the end projecting out. The flask was placed over a furnace, **B**, and a jar of air, **D**, was placed over the projecting end.

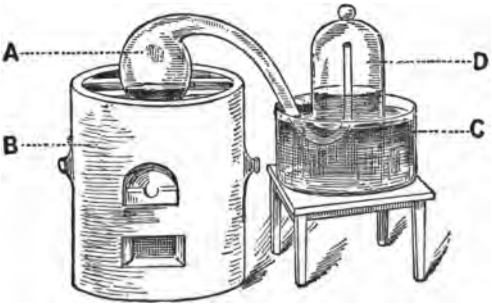

FIG. 9

The air in the jar, then, was in communication with the mercury in the flask, but shut off from the outer air.

The flask was heated for twelve days, nearly to the boiling point of mercury, and during this time the mercury inside was seen to be gradually covered with dark red scales (the calx of mercury). The volume of the air inside the jar was observed to be slowly diminishing. When the volume of air decreased no more the apparatus was allowed to cool.

The original volume of air in the jar was 50 *cubic inches.* It was found to have diminished by 8 *cubic inches.* The red calx was collected. It was placed in a small retort, and heated over a furnace, and the gas given off was collected in a jar. *Eight cubic inches* were obtained—just the amount, in fact, which had previously disappeared. Moreover, when Lavoisier mixed this 8 cubic inches with the 42 cubic inches left in the jar before, **he obtained a mixture which behaved precisely as ordinary air.**

Thus by 1775 it was clearly established that air contains two parts—one **active**, and about one-fifth by volume, the other **inactive**, and about four-fifths by volume.

CHAPTER VI

PREPARATION OF ACTIVE AIR

ALTHOUGH both red lead and calx of mercury give active air on being heated, yet it is difficult to obtain large quantities of the gas from these substances.

Potassium chlorate (a white crystalline salt) is found to melt and evolve this gas when heated, but by mixing it with various bodies, such as **sand** or **manganese dioxide**, the gas is evolved more easily, and at a lower temperature, the added material remaining unchanged at the end of the operation.

Exp. 26.—To prepare active air from potassium chlorate and manganese dioxide.

Take a large test-tube, about 15 × 2·5 cm. in size, and fit it with a softened cork, through which a hole is bored (Fig. 10).

Pass a short piece of glass tubing about 10 cm. long through the hole, and connect it by rubber tubing with a long **delivery tube**, about 60 cm. in length.

Fix the test-tube horizontally between the rings of a retort stand, at such a height that the delivery tube dips into a sink or trough of water.

Support over the end of the delivery tube a jar full of water, dipping mouth downwards into the trough.

This jar can either be held in position or stood on a "**beehive shelf,**" as in the diagram.

Place in the test-tube a mixture of potassium chlorate (4 parts) and manganese dioxide (1 part), and connect up the apparatus.

Heat the test-tube gently with a bunsen flame. Bubbles of air are first given off at the end of the delivery tube, because of the expansion caused by the heat. Allow these to escape. Then the active air is allowed to enter the jar, and it displaces the water inside. This method of collecting gases was invented by **Priestley**, and is called "**collection over water**," or "**collection by displacement of water.**" He also called the trough a "**pneumatic trough**," because it is used for the collection of airs or gases (Greek *pneuma* = air).

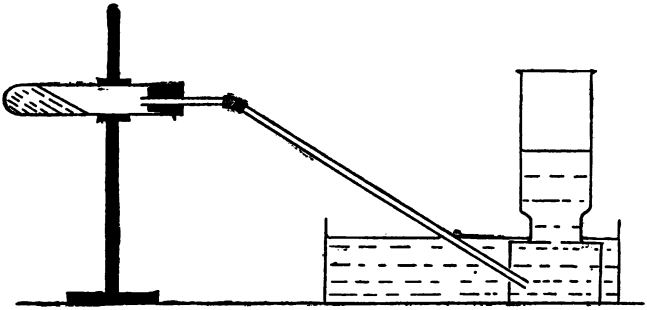

FIG. 10

As each jar is filled with gas, place over its mouth a ground-glass plate, and remove it from the trough.

Place a fresh jar filled with water over the end of the delivery tube, and continue the process till several jars of gas are ready. Notice that, at first, the gas is cloudy. This is due to fine particles of solid carried over from the test-tube.

When enough gas is collected remove the delivery tube from the water, and allow the apparatus to cool.

Notice that the gas is *colourless*, *odourless* and *tasteless*, and *evidently not very soluble in water.*

Experiments with Active Air.—We have already found that a glowing splint is rekindled in this gas. Let us now try the effect of introducing into it various combustible bodies, and let us examine more closely the calces or fumes produced. For this purpose we may use **litmus**, a blue colouring matter obtained from a lichen, *which we have seen has the property of being turned red by acids.*

Exp. 27.—Take a little litmus solution in a test-tube, and add a drop of **sulphuric acid** from your bench. Notice that a bright red tint is produced. Try a few drops of the other acids on your bench in the same way, and see if they also affect litmus.

Exp. 28.—To burn phosphorus in active air.

Take a wire stand, such as you used in Exp. 11, and holding it upside down, place on the spiral a fragment of **phosphorus.** Just light the phosphorus by holding it in a bunsen flame for a second, and then plunge it into a jar of the gas (Fig. 10).

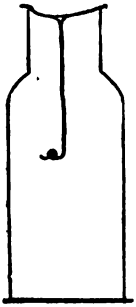

FIG. 11

The phosphorus burns with a very brilliant white light, evolving clouds of fumes. Keep the glass plate over the jar to prevent them escaping into the air.

When the burning has ceased, take out the stand, pour a little water into the jar, and shake up. Presently the fumes dissolve in the water. Now pour in a little **blue litmus solution.** Observe that it *immediately turns red.* Evidently the *solution of phosphorus fumes is acid.*

Exp. 29.—To burn sulphur and carbon in active air.

Repeat Exp. 28, using a small piece of **sulphur**, which burns with a faint blue flame in air. Notice that it burns much more vigorously in active air, and that the fumes evolved have a very *suffocating odour.* Add water and shake up to dissolve these fumes. Repeat Exp. 28 again, using a piece of charcoal, which merely glows when heated in air. Notice that it sparkles and glows intensely when burning in active air. Add water and shake up as before.

Now pour a little **blue litmus** solution into each jar. Notice that the *sulphur fumes give a strong acid solution*, while those from the charcoal only tinge the litmus a *wine-red colour.*

Now **Lavoisier** observed that such bodies as sulphur burn in active air and form fumes which give acid solutions with water. He therefore proposed the name **oxygen** for this gas, deriving it from two Greek words meaning "**acid**" and "**to produce.**" He believed that all acids contain oxygen as an essential constituent. This name is still retained. From it we obtain a name for the various calces and fumes we have been dealing with, viz. **oxide**; thus litharge is *oxide of lead*, calx of mercury is *oxide of mercury*, iron rust is *oxide of iron*, and the white fumes of phosphorus are *oxide of phosphorus.*

To find if all oxides give acid solutions in water.

Exp. 30.—Take a strip of **magnesium** ribbon about 15 cm. long. Ignite one end and introduce it into a jar of oxygen. It burns even more brightly than in air, while *magnesium oxide* is formed. Pour in a little water and shake up thoroughly. Now add a little litmus solution which has been turned a very faint pink by a drop of acid. The colour turns to a *pale blue*. Here then is an oxide which is certainly not acid, but which has precisely the opposite action on litmus.

Exp. 31.—Take a scrap of **sodium**, a *soft*, *white metal* which tarnishes or **oxidises** so rapidly in air that it has to be kept under **naphtha**—a liquid containing no oxygen. Place it on a *deflagrating spoon* and heat till it begins to burn, then put the spoon into a jar of oxygen. The sodium burns with a clear yellow flame, leaving behind a white residue of *sodium oxide*.

When cool place a little water in the spoon. The oxide rapidly dissolves. Pour the solution into a test-tube containing a little reddened litmus. Notice that the *litmus turns blue*.

This is another example of an oxide which does not give an acid solution with water.

Exp. 32.—Make a spiral of fine **iron wire**. Fix a little piece of wax vesta at the end of the spiral, and attach the whole to a wire stand.

Light the vesta, and plunge the stand into a jar of oxygen. The wax burns brightly, and ignites the iron, which sparkles brilliantly, drops of molten oxide falling to the bottom of the jar, which should contain a little water. Take out a piece of the oxide, and place it on a moistened litmus paper. No effect is seen. *The oxide is neutral to litmus.*

EXERCISE

Try the effect of litharge, lime, copper oxide and caustic potash on red and blue litmus.

CHAPTER VII

DIFFERENT KINDS OF OXIDES

We have seen that some oxides dissolve in water and form acids, turning litmus red, while others have the opposite effect, turning it blue. Others again have no effect on litmus.

Now substances which act on litmus, turning it blue, have been known for a long time. The ash obtained when wood is burned, acts in this way. **Washing soda** and **lime** have similar effects.

The term **alkalies** is applied to these bodies, having been first given by the Arabian chemists to wood ashes.

Exp. 33.—To find the effect of an alkali on an acid.

Place a piece of **caustic soda** in an evaporating dish. Notice that it feels *slimy* and appears to act upon the surface of the skin, hence its name caustic (=burning). Avoid handling it overmuch. It is seen to become *moist* presently. It has such an attraction for water that it absorbs moisture from the air. Such substances are called **deliquescent** (=becoming liquid). Cover it with water, noticing that it soon dissolves and the water becomes warm. Cautiously taste a drop of the solution. Pour in a little solution of litmus which has been *reddened* by acid. The litmus at once turns *blue*. Caustic soda is evidently **alkaline**.

Take the bottle of dilute **hydrochloric acid** from your bench and add acid drop by drop to the basin, stirring the while with a glass rod, till the litmus turns a *purple or neutral tint*.

What has happened to the acid? Is it possible that both acid and alkali are still present as such, though the action of each on the litmus is neutralised by the other? Can we find if this is the case? Suppose we drive off the water and see if we obtain the caustic soda as residue.

Place the dish on a tripod stand and heat carefully with a bunsen flame (Fig. 12). The water boils away, and presently a white residue remains. Is this the acid or the alkali we started with?

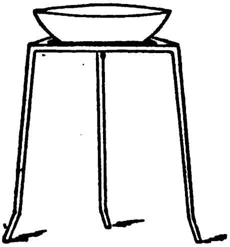

FIG. 12

Does it become *moist* when exposed to the air? Taste a little on the tip of your finger. Does it taste the same? It will be found to have rather a *salty or saline taste*, and it seems likely that we have formed a new substance.

A substance thus made by the **neutralisation** of an acid by an alkali is termed a **salt.**

The salt we have just made is ordinary *common or table salt.*

EXERCISE

Prepare salts as above from caustic potash, neutralised by (1) dilute sulphuric acid and (2) dilute nitric acid.

Can we make a salt by using such an oxide as litharge and an acid? In other words, Can we neutralise an acid with litharge, or some such insoluble oxide? Evidently litmus would not be of much use to indicate when the solution is neutral, because the litharge would make the liquid turbid. How can we be sure that the acid will be neutralised? Could we not use an *excess* of the oxide to be quite certain that all the acid will be used up?

Exp 34.—To try to neutralise nitric acid by litharge.

Place about 10 c.c. dilute **nitric acid** in a test-tube. Add a little **litharge** and warm. It soon dissolves. Keep on adding the oxide till no more dissolves, even on heating, and a residue of litharge is left. Hence the acid will take up no more and seems to be *neutralised.*

Fold a filter paper into half and again into a quarter. Open it out so that a cone is formed. Fit the cone into a funnel. Support the funnel in a ring of the retort stand, and place an evaporating basin below it (Fig. 13).

Add an equal volume of water to the contents of the test-tube, warm and pour on to the filter paper. A clear liquid runs through, leaving the sediment on the paper.

This process of separating a liquid from an insoluble solid is called **filtration.** The clear liquid is called the **filtrate** and the sediment the **residue.**

Place the evaporating basin on a tripod stand and drive off the water by heat.

Notice again that a white residue is left, which is soluble in water and neutral to litmus.

A salt has been formed.

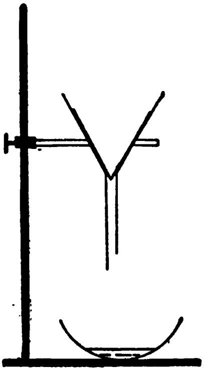

FIG. 13

EXERCISE

Form salts by neutralising the acids on your bench with copper oxide.

All oxides, soluble or insoluble in water, which will neutralise acids to form salts are called bases or basic oxides. Those bases which dissolve in water and turn red litmus blue are called **alkalies.**

OXIDES.

Acidic.	Basic.	
Dissolve in water to form acids.[1]	Usually oxides of metals. Neutralise acids to form salts.	
	Soluble in water.	Insoluble in water.
e.g. Oxide of phosphorus. Oxide of sulphur.	Alkalies. *e.g.* Oxide of sodium.	*e.g.* Oxide of lead.
Acids turn blue litmus red.	Alkalies turn red litmus blue.	No effect on litmus.
Acid + Base = Salt.		

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[1] Hence acidic oxides are often called *Anhydrides*, meaning acids deprived of water.

Salts are named so as to show from what base and acid they are formed.

Salts of sulphuric acid are called **sulphates.**
,, nitric ,, **nitrates.**
,, hydrochloric ,, **chlorides.**
,, carbonic ,, **carbonates.**

For instance, litharge and nitric acid form lead nitrate, copper oxide and sulphuric acid form copper sulphate.

CHAPTER VIII

INACTIVE AIR

In our experiments on burning we found there always remained *four-fifths of air* unused, and that this gas would not support combustion.

Let us now examine this **inactive air** more fully. We can evidently prepare it by taking oxygen from atmospheric air by means of such substances as **phosphorus**, **copper**, etc.

This gas was first studied by **Lavoisier**, who gave it the name **azote**, a name implying inability to support life. But because it was found in nitre it received its present name of **nitrogen.** In France it is still called azote.

Exp. 35.—To prepare nitrogen and examine its properties.

Take a hard glass tube (**A**, Fig. 14) 50 × 2 cm. in size. Round off the ends and fit it with two bored corks each provided with short pieces of glass tubing. Pack it with copper turnings.

To one end attach by rubber tubing a delivery tube dipping into a pneumatic trough containing a beehive shelf and collecting jar full of water, **C**.

Connect the other end to a large bottle, **B**, into which water is poured by means of a tube connected to the water tap. The air from the bottle is thus slowly forced over the copper. The tube is placed on a furnace and the copper heated red hot (Fig. 14).

When the tube is hot, water is allowed to slowly enter the bottle, causing a current of air to pass over the copper. This deprives the air of its oxygen. Copper oxide is formed and nitrogen is collected in the glass jar.

Obtain several jars of the gas.

Notice that it is a *colourless*, *tasteless* and *odourless* gas. It cannot be very *soluble in water*, since it was collected by displacement of water.

Plunge a lighted taper into a jar of nitrogen.

The taper is immediately extinguished. Try with lighted phosphorus.

Shake up the gas with a little litmus solution—there is no effect. *The gas is neutral* and *does not burn or support combustion.*

In fact, nitrogen is so inactive that it is very difficult to get any substance to combine with it.

Its presence in the air serves to *dilute* the oxygen. Animals require oxygen to breathe, just as a taper requires it to burn in.

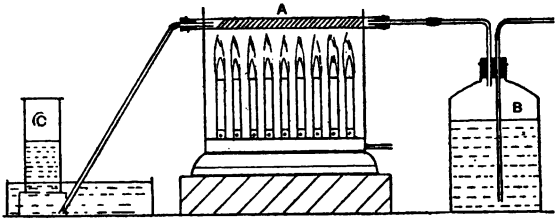

FIG. 14

Pure oxygen would cause a too rapid combustion, so the nitrogen serves to modify its action.

Yet when the lungs are too weak to fulfil their proper function in extracting oxygen from the air, as in some diseases of the chest, pure oxygen is frequently administered.

CHAPTER IX

INFLAMMABLE AIR

IN 1766, **Cavendish**, a chemist and physicist of great eminence, discovered that by the effect of certain metals on acids an *inflammable gas* was given off.

Exp. 36.—To find the action of zinc on acids.

Take a piece of granulated **zinc** in a test-tube, and add to it a few c.c. of **dilute sulphuric acid**. A rapid effervescence of gas takes place. Apply a light to the mouth of the tube. Notice that a slight explosion takes place, but that when the gas has driven out all the air it burns with a *faint blue flame*. The test-tube becomes warm by reason of the chemical action which is going on inside.

Add more zinc till the effervescence ceases. This means that the acid is used up. What has been formed besides this inflammable air? Add water to the test-tube, warm, and filter into an evaporating basin. Place the dish on a tripod stand, and evaporate to dryness.

A white mass remains. Does it not appear that we have neutralised the acid by the metal, forming a **salt**, viz. **zinc sulphate?**

EXERCISES

1. Compare this substance with zinc sulphate obtained by neutralising sulphuric acid with zinc oxide.

2. Try in a similar manner if inflammable air is evolved from the action of the following metals on the acids on your bench: tin, iron filings, magnesium ribbon, aluminium foil. Obtain salts where possible.

Exp. 37.—To collect inflammable air, and observe its properties.

Take a flask holding about 500 c.c. Fit it with a softened cork, through which two holes have been bored. Pass a *thistle funnel* through one hole, and a right-angled bend through the other.

Place granulated zinc in the flask, sliding it in carefully to avoid fracture. Cover the zinc with water, and take care that the thistle funnel dips below the surface of the water.

By means of rubber tubing connect the right-angled bend with a delivery tube, which passes under a beehive shelf placed in a pneumatic trough containing water (Fig. 15).

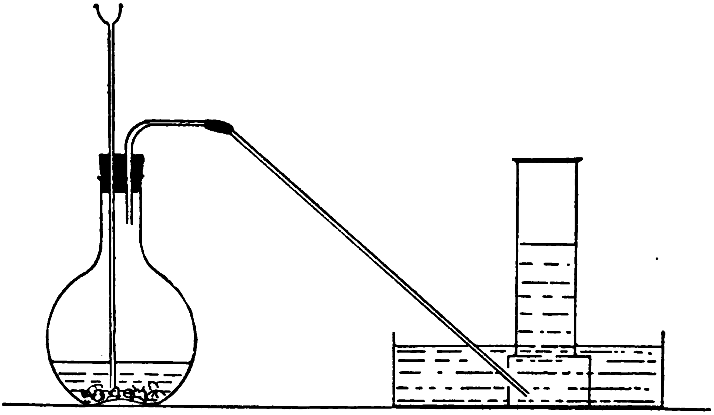

FIG. 15

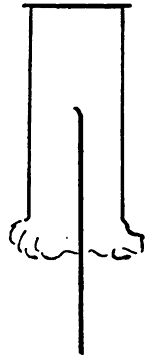

FIG. 16

Pour strong **sulphuric acid** down the thistle funnel till a brisk effervescence starts. The first bubbles of gas which issue from the delivery tube will be mixed with air, and therefore explosive. Collect the gas first of all in a test-tube full of water, and see if it burns quietly. If not, continue testing in this way, till the gas burns without explosion.

Then place a jar full of water on the beehive shelf and collect by displacement of water, as in the case of oxygen. Obtain several jars of gas in this way, noticing that the inflammable air thus obtained is *colourless*, *tasteless*, *odourless* and *practically insoluble in water.*

Exp 38.—Experiments with inflammable air.

Hold a jar in an *inverted* position and take off the cover. Thrust a lighted taper up into the gas. The taper is extinguished, but the gas burns at the mouth with a pale blue flame. Reignite the taper at the burning gas, and again thrust it up into the jar. Evidently the *gas burns in air*, but *does not support combustion.*

Exp. 39.—A. Stand a jar of the gas on the bench and take off the cover. After a few seconds apply a lighted taper. A *loud explosion* is heard. The gas seems to have risen partly from the jar, air taking its place, forming an explosive mixture. Evidently inflammable air is *very light gas.*

B. Uncover another jar of gas, but this time place it under a jar of air (Fig. 17). After a few seconds apply a lighted taper to the jars.

Each jar contains an explosive mixture of air and inflammable air.

C. Allow a third jar to stand open, but inverted. After the same lapse of time as above bring near a lighted taper. There is no explosion and the gas burns quietly.

Exp. 40.—Support a large inverted beaker from the arm of a balance. Counterpoise it carefully. Hold beneath it a jar of the gas. The gas rises into the beaker and so pushes it up. Show that the beaker now contains inflammable air by applying a light.

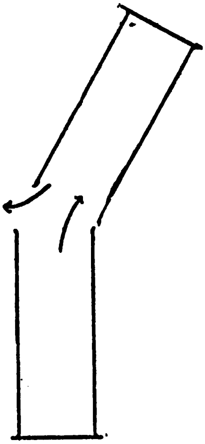

FIG. 17

This property of *extreme lightness* has been utilised in the filling of balloons.

We may take advantage of the gas being light by collecting it by upward displacement of air. In this method the gas is led up by a delivery tube, into an inverted gas jar. Being light it tends to rise to the top of the jar, thus displacing the air.

This method has the advantage of obtaining the gas fairly dry (Fig. 18).

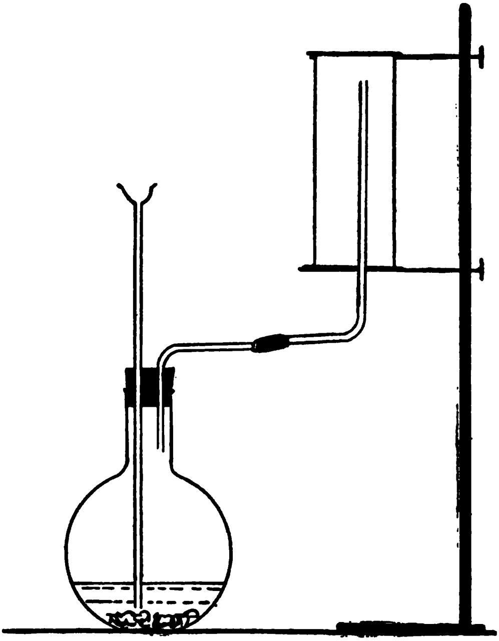

FIG. 18

Source of Inflammable Air.—Has this gas come from the metal or from the acid? We may obtain it from zinc and either sulphuric or hydrochloric acid, or from magnesium, iron and some other metals when acting on either of these acids. In the case of zinc, we found that a salt of zinc remained, similarly salts of the other metals would be formed. Hence it seems as though the acid were the source of the gas.

EXERCISE

Think how an acid could be made (*e.g.* from sulphur), and so try to discover whether inflammable air is a constituent of water.

CHAPTER X

COMPOSITION OF WATER

Exp. 41.—To find what is formed when inflammable air burns.

Take a flask containing zinc and water. Fit it with cork, thistle funnel, and right-angled bend, as in Exp. 37. By means of rubber tubing attach a jet to the end of the right-angled bend. Pour acid down the funnel and so start the evolution of gas. When the gas issues in a pure condition, light it at the jet. Hold over the flame a clean, dry, gas jar. Notice that a dew is immediately deposited. Throw in a little anhydrous copper sulphate. *It at once turns blue.* It seems as though water were formed.

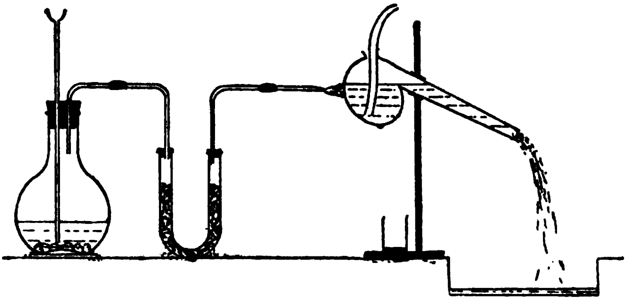

FIG. 19

But this water may have been carried over from the flask. To make sure that this shall not happen, attach the jet to a U-tube containing **calcium chloride**, and connect this to the right-angled bend in the flask. The gas passing over the calcium chloride will be dried, because this substance is *extremely deliquescent.*

Fit a retort in a stand, and place in the *tubulure* a rubber tube connected with the water tap. Allow water to flow into the retort, escaping by the neck (Fig. 19). We can thus hold the flame against the cool sides of the retort, and if water be really formed, it will be condensed there. Light the gas at the jet, first testing to see if it is free from air. Allow the flame to play on the retort. Drops of liquid collect and drop into a beaker placed below. The liquid is *colourless* and *mobile.* Is it water? Now, it has been found that a pure substance can always be recognised by certain of its physical properties, such as *density, boiling point* and *freezing point.*

Therefore when sufficient of the liquid is collected, find, for instance, its boiling point and density. Also add to it some anhydrous copper sulphate.

By these tests it may be proved to be water. Water, then, seems to contain as one constituent **inflammable air**, which is combined with the **oxygen** of the air.

Cavendish first showed that water contained these gases by exploding in a strong glass globe inflammable air and ordinary air in various proportions. When mixed together in certain quantities, viz. 423 measures of inflammable air to 1,000 of common air, he observed that all the inflammable air vanished, together with one-fifth of the common air, dew being formed in the globe.

But he had found that ordinary air contained about 21 per cent. of oxygen, hence this experiment indicated that 200 volumes inflammable air and about 100 volumes of oxygen would completely combine to form water.

EXERCISE

Devise an experiment to show that when inflammable air burns in a closed space of air one-fifth of the air disappears (see Fig. 20).

Can we form water in some other way from oxygen and this inflammable air? What substance contains oxygen in an available form? Why not try passing the gas over heated **copper oxide**, for instance? The copper oxide, as we know, may be obtained by heating copper in air, thus containing oxygen, which may possibly unite with the inflammable air.

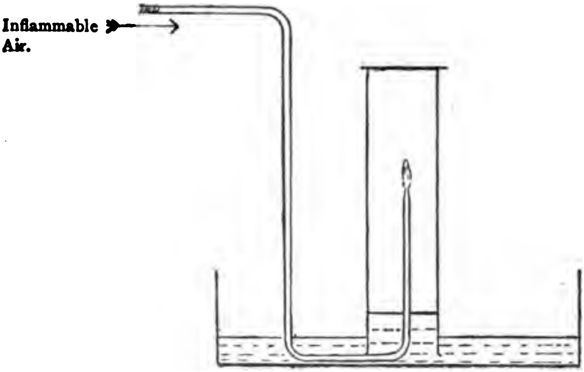

FIG. 20

Exp. 42.—To find if water is formed when inflammable air passes over hot copper oxide.

Take a flask containing granulated zinc and water fitted for the evolution of inflammable air and connected to a calcium chloride drying-tube.

Take a reduction bulb containing copper oxide, and attach it by rubber tubing to the drying tube (Fig. 21).

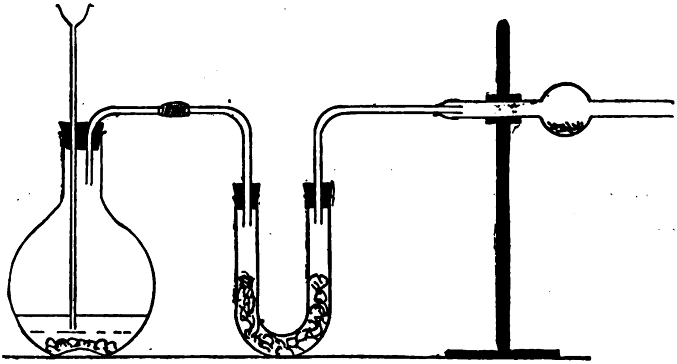

FIG. 21

Pour strong sulphuric acid into the flask and allow a steady stream of inflammable air to pass over the copper oxide. Test the gas, and when it is pure begin to heat the reduction bulb. Presently drops of liquid collect on the cool sides of the bulb-tube. Place in a little anhydrous copper sulphate. It is at once turned blue. Water is evidently formed. Notice that the copper oxide is now replaced by *metallic copper.* An action of this sort, in which a substance is deprived of oxygen, is called **reduction.** The inflammable air has reduced the copper oxide to copper, and water has been formed.

We have some reason now for thinking that water contains oxygen. If we acted on water with some substance which has a great attraction or affinity for oxygen, should we not obtain an oxide and inflammable air? In other words, should we not **analyse** (=to loose asunder) water? So far we have been forming water by putting together the two substances, oxygen and inflammable air, *i.e.* we have formed water by **synthesis** (= to put together).

Exp. 48.—To analyse water by red-hot iron.

Pack a piece of gun barrel or *combustion tubing* about 40 cms. long with iron nails. Fit each end of the barrel with bored corks carrying short pieces of glass tubing. Connect one end with a flask containing water, which can be heated, thus driving a current of steam over the iron. To the other end attach a delivery tube, dipping under a beehive shelf placed in a pneumatic trough containing water. The flask is provided with a long, straight tube dipping under the water. This acts as a *safety tube.*

Heat the gun barrel to bright redness in a furnace (Fig 22).

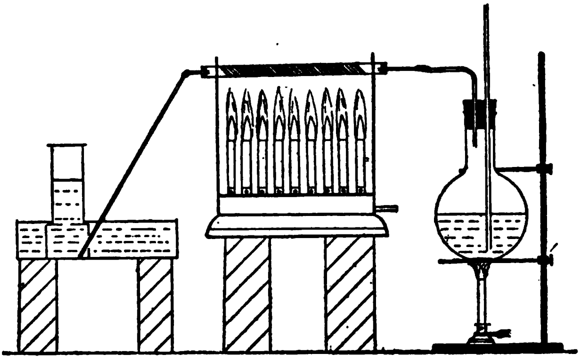

FIG. 22

Then heat the flask till the water boils. Connect up so as to allow steam to pass over the red-hot iron. Bubbles of displaced air come oft first. Allow these to escape and then collect the issuing gas in a jar full of water standing on the beehive shelf.

Apply a light to the gas and notice that it burns with the pale blue flame characteristic of inflammable air.

Examine the residue in the gun barrel.

This experiment was first performed by **Lavoisier,** and because he thus showed inflammable air to be a constituent of water he gave to it the name **hydrogen,** derived from two Greek words and meaning **water producer.**

To Analyse Water by other Metals.—Many metals act on water, forming metallic oxides, and liberating hydrogen.

Exp. 44.—Wrap a fragment of **sodium** in a piece of copper gauze, and drop it into a trough of water. The sodium acts on the water quite vigorously, and bubbles of gas escape. Collect these in a gas jar full of water held over the metal (Fig. 23). Test the gas by holding a lighted taper to the mouth of the jar. Notice that the characteristic blue flame of hydrogen is formed. Add a little red litmus solution to the water in the trough.

The litmus turns blue, because the sodium oxide formed is alkaline (see p. 121).

Place a small piece of **potassium**—a metal very similar to sodium—on water in a trough. The metal floats on the surface, and, so actively does it take up oxygen from the water, that the evolved hydrogen catches fire and burns with a flame coloured violet by the potassium. Again notice that an alkaline solution of potassium oxide is formed.

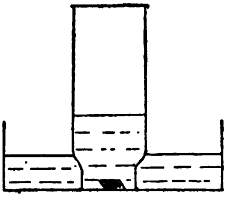

FIG. 23

Exp. 45.—To analyse water by magnesium.

Take a large flask containing water. Fit it with a cork, through which two holes are bored. A long, straight glass tube passes through one hole nearly to the bottom of the flask; this acts as a safety tube. Through the other hole is a right-angled bend, connected to a small flask which serves as a trap for any water which may be condensed. A right-angled bend passes from this flask to a reduction bulb, containing magnesium (Fig. 24).

Heat the large flask till the water boils vigorously, gently warming the reduction bulb the while to prevent any steam condensing on it. When a brisk current of steam issues from the end strongly heat the bulb containing the magnesium. Presently the metal catches fire, and then hydrogen can be ignited at the end of the bulb-tube.

Notice the substance left in the bulb. Does it not resemble the body obtained when magnesium is burnt in air?

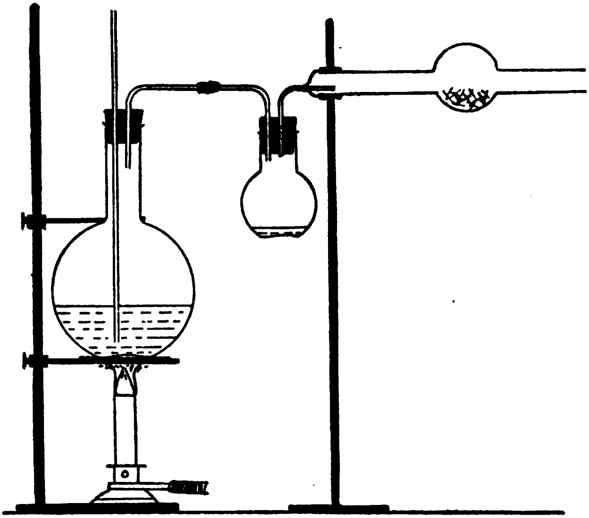

FIG. 24

We may express the action of metals on water in the form of **equations.** Thus :—

Iron + water yield iron oxide + hydrogen,
Sodium + water „ sodium oxide + hydrogen,
Potassium + water „ potassium oxide + hydrogen,
Magnesium + water „ magnesium oxide + hydrogen,

and generally

Metal + water „ metallic oxide + hydrogen.

Copper, gold, and platinum have no action on water, even if heated to a very high temperature.

CHAPTER XI

WATER

We have seen in the preceding chapters how water may be analysed and synthesised, and how it may be recognised by means of its physical constants, such as freezing point, boiling point and density. We must now examine its properties more fully.

When pure, water is a **colourless, tasteless, odourless, mobile liquid.** In large quantities it has a bluish green tint. The mass of 1 c.c. of pure water at 4° C. is taken as the standard of mass, and is called **1 gram**; hence the density of water at this temperature is unity.

Under normal atmospheric pressure (760 mm. of mercury) it boils at 100° C. and freezes at 0° C., but these temperatures vary if the pressure be altered.

Exp. 46.—To show the effect of reduction of pressure on the boiling point of water.

Boil some water in a flask, which can be closed with a well-fitting cork. Steam is evolved, which, after a time, drives out most of the air from the flask. Remove the flame and cork the flask tightly. Cool by holding the flask under a stream of water. Notice that the water begins to boil, although its temperature is under 100° C. The atmosphere of steam which was above it has now been condensed, thus a great reduction of pressure has taken place, resulting in the water boiling at a lower temperature.

In a steam boiler, where the pressure is very great, the boiling point of water is raised considerably. In high altitudes, where the atmospheric pressure is small, cooking by means of boiling water in the ordinary way is impossible, because the water is not sufficiently hot. Vessels are used, therefore, which can be screwed up, and in which the pressure of the steam can be raised high enough to enable the requisite temperature to be attained.

Effect of Heat on Water.—If heat be applied to ice, which has been cooled to a temperature below 0° C., it follows the usual behaviour of solids and expands. When 0° C. is reached the temperature remains steady till all the ice is melted, during which process **much heat is absorbed or rendered latent.** One gram of ice at 0° C. requires nearly 80 *units of heat* to bring it into the condition of water at that temperature. At the same time every 10 volumes of ice contract to 9 volumes of water.

On continuing the heating, the volume of the water becomes less and less till the temperature reaches 4° C. **At this temperature a given mass of water has its least volume and therefore its maximum density.**

From 4° C. the water gradually expands, as the heating goes on, till the temperature is 100° C., at which temperature the thermometer remains constant till all the water is converted into 1728 times its volume of steam—the pressure being 760 mm. 537 *units of heat* are required to turn 1 gram of water at 100° C. into 1 gram of steam at that temperature.

And, finally, as the heating goes still further, the steam expands at a greater rate, behaving like a gas.

These changes can be expressed in diagrammatic form, as is seen below.

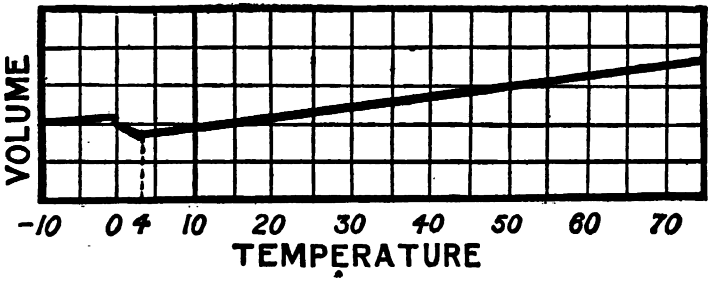

FIG. 25

Two straight lines at right angles are taken as axes. Let the vertical line be marked off to represent volumes and the horizontal line to represent temperatures. Ordinates are erected at each temperature to correspond with the volume of the water or ice at that temperature and so a curve can be obtained showing the relation between these quantities.

To show the point of Maximum Density of Water.—A tall jar is taken, through the side of which are fixed by means of corks two thermometers. Around its middle is a bath which contains ice (Fig. 26). The jar is filled with water, and the temperatures indicated by the thermometers are read from minute to minute. **The mercury in the lower thermometer gradually sinks to 4° C.**, for as the water cools, as is shown by the upper thermometer, it becomes more dense and sinks to the bottom. Presently there is a uniform temperature of 4° C. As the temperature of the water at the surface falls below this the water now becomes less dense, and consequently does not sink to the bottom, but remains at the top and gets colder and colder till it finally freezes. *Still the lower thermometer indicates* 4° C. This fact is of great importance in nature. For did water steadily contract as it cooled down to zero, then the coldest water would always be found at the bottom, so that ponds, rivers, and seas would freeze there first and then congeal to a solid mass, destroying all their living occupants. And, moreover, the heat of summer would not serve to thaw this accumulation of ice.

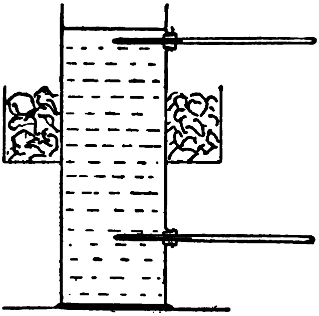

FIG. 26

How Water is Heated.—Liquids do not convey heat in the same way as solids. An iron poker placed in a fire conveys heat by **conduction,** *i.e.* the heat passes from particle to particle of the iron.

But in a liquid the heat is conveyed principally by the liquid itself in **convection currents,** which are due to the difference of density between the hotter and colder parts.

Exp. 47.—To show the convection currents in water.

Place some sawdust in a large beaker full of water. Get the sawdust uniformly diffused through the water. Support the beaker on a stand and heat with a bunsen flame.

As the heating proceeds, notice that the sawdust is carried about with the convection currents of the water. Frequently a complete circulation can be obtained in this way, the sawdust rising up the middle of the beaker, and descending again by the sides.

This **circulation** of a liquid is also to be observed in Exp. (p. 136), the cooler, denser water sinking to the bottom till a temperature of 4° C. is attained.

Buildings are frequently heated by utilising this property. Water is heated in a boiler in the basement, and then rises through a series of pipes passing through the rooms to a tank placed at the top of the building. As it cools it becomes more dense, and finally at the top of the building it flows down to the boiler by another set of pipes. A constant circulation is thus set up.

Water is a *bad conductor of heat.* Heat travels through it very slowly if convection currents are prevented. If a piece of ice be loaded with wire and sunk to the bottom of a test-tube full of water, which is heated near the top, the surface water may be boiled without causing the ice to melt.

CHAPTER XII

SOLUTION

We have already had several cases in which water was used to dissolve a substance. In such cases a **solution** in water was formed, water being the **solvent**, and the dissolved substance assumed the liquid state.

Water has so great an action as a solvent that it is a matter of extreme difficulty to obtain it in a perfectly pure state.

The purest form of naturally occurring water is *rain.* As the rain penetrates the surface of the earth it takes up a considerable amount of matter, depending on the nature of the ground it passes through.

More impurities are added when the rain becomes part of a river, as *organic matter* is frequently found in it, being obtained, for instance, from sewage, and by the time the sea is reached a large amount of substance is held in solution.

In the **Clyde**, above Glasgow, *e.g.*, the dissolved solids form about ·01 per cent.; in the **sea water** they average about 3·6 per cent.; but in the **Dead Sea**, where pure water is rapidly lost by evaporation, the percentage rises to 24.

Exp. 48.—To show that tap water contains dissolved solids.

Evaporate to dryness a few c.c. of tap water in a watch-glass placed on a wire gauze on a tripod stand. Notice a grey deposit of solid matter left behind.

If a larger weighed amount of water had been evaporated to dryness in a counterpoised dish we could have found the percentage of solid matter in the tap-water.

Exp. 49.—To show that tap water contains dissolved gas.

Take a flask holding about 500 c.c. Fit it with a bored cork through which passes a right-angled bend, connected by rubber tubing with a delivery tube. Let the delivery tube dip under water in a pneumatic trough. Stand over the end of the tube a gas jar full of water (Fig. 27).

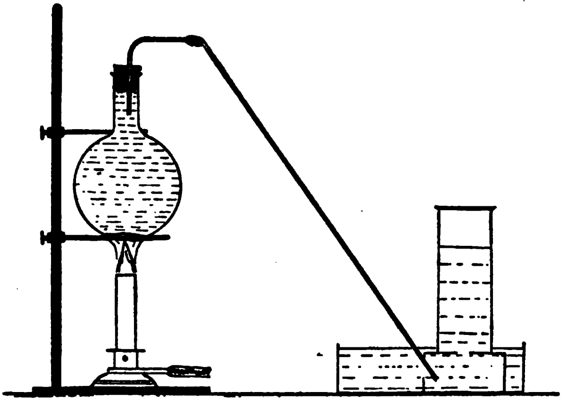

FIG. 27

Fill up the flask with tap water, and press in the cork. This will fill the bent tube with water.

Suck up water in the delivery tube, and, pinching the rubber tube to prevent the water escaping, join up with the right-angled bend. If any air remains, hold the delivery tube upright, and tap it till the bubble of air is driven out.

Place the flask on a retort stand and heat. Bubbles of air are seen to be driven off and are collected in the gas jar. Presently the water boils, and the steam is heard condensing in the trough with a rattling noise. Now place a glass plate over the mouth of the jar, and remove it from the trough. Only a small amount of air has been collected. We could not measure this accurately by using an ordinary measuring jar. Therefore take a burette containing water. Read off the volume of water in the burette, and run water from it into the jar till the latter is completely filled. The volume of water delivered from the burette gives the amount of air driven off. Now measure the volume of the flask by pouring the water it contains when filled up to the cork into a measuring jar. Thus find the percentage of air contained in tap water.

EXAMPLE

First reading of burette = 50 c.c.
Second reading of burette = 40·5 c.c.
∴ Volume of air driven off = 9·5 c.c.
Volume of flask = 330 c.c.

$$\therefore \text{Percentage of air in the water} = \frac{9\cdot5 \times 100}{330} = 2\cdot9.$$

Tap water, then, contains both *dissolved air* and *dissolved solid matter*. How shall we obtain a specimen of pure water from it? If the water is heated and turned into steam, and then the steam condensed, we ought to obtain the water free from the solids certainly, and partially free from the dissolved air, although on cooling the condensed water will take up air from the surrounding atmosphere. This process of turning a liquid into its vapour, and then condensing the vapour, is called **distillation**.

Exp. 50.—To purify tap water by distillation.

Fit up a retort in a stand, as in Fig. 28. Place on the neck of the retort a small flask to act as receiver. Keep the receiver cool in a trough of water. Place water in the retort till it is about one-third full, and fit the tubulure of the retort with a cork.

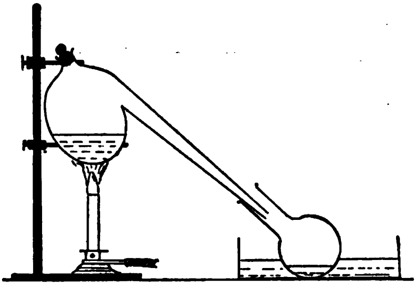

FIG. 28

Heat the retort gently. The temperature rises to about 100° C., and then the water boils. The evolved steam begins to condense on the neck of the retort and runs down into the receiver. This distilled liquid is usually termed the **distillate**. Rinse out the receiver with the first portion of the distillate.

Test the distilled water by evaporating a little to dryness, as in Exp. 48. No residue is noticed. Taste a little. It tastes "flat" and rather unpleasant. The presence of dissolved gas gives water its usual taste. Add litmus, and notice that pure water is neutral. Boil a little of the distilled water in a test-tube. Hold a thermometer in the steam, just above the boiling liquid. Notice that the mercury stands at about 100° C. Take the barometer reading.

Solubility.—It is a matter of common experience that water will not dissolve an unlimited quantity of a soluble substance, such as sugar, for instance. When no more will dissolve we have a **saturated solution**. Again, hot water is usually found to dissolve more of a soluble substance than cold water, so that the **amount dissolved to form a saturated solution varies with the temperature of the solvent.**

Thus we may define the **solubility** of a substance in water at any particular temperature to be the **weight in grams of the substance required to saturate 100 gm. of water at that temperature.**

EXERCISE

Weigh into a beaker 100 gm. of water. Weigh on a piece of paper 100 gm. of finely powdered **nitre**. Add this bit by bit to the water, stirring continually till no more will dissolve. Find the weight of the remaining nitre. Find the temperature of the solution. Thus obtain the solubility of nitre at that temperature.

Repeat this experiment, using sugar, salt, sal-ammoniac, etc.

By taking axes and measuring solubilities on one and temperatures on the other, we can make a curve, showing for any particular substance the relation between solubility and temperature. This is called the **solubility curve** of the substance.

Exp. 51.—To make a solubility curve for potassium chlorate.

Let each set of students work at a separate temperature at which to find the solubility of **potassium chlorate**. Thus a large number of experiments can be made, giving plenty of points from which to obtain the curve.

Weigh out a clean, dry evaporating dish. Take a large test-tube half full of water and make a *boiling*, *saturated solution* of potassium chlorate. Meanwhile take a beaker full of water containing a thermometer, and get this approximately to the temperature at which you are working. Place the test-tube in the beaker and shake it continually. Then allow the separated crystals of the potassium chlorate to settle, and quickly take the temperature of the saturated solution. Carefully decant off the clear liquid into the weighed dish, avoiding any transference of solid, and weigh it again. This gives the weight of solution used.

Place the dish on a wire gauze on a tripod stand and heat carefully to drive off the water. When the solution gets pasty, heat very gently to *avoid loss by spurting*. When all the water has been driven off, allow the dish to cool and weigh again. This gives the weight of solid. Insert your weighings in your note-book thus :—

Wt. of dish = 28·17 gm.
Wt. of dish + solution = 37·2 gm.
∴ Wt. of solution = 9·03 gm.
Wt. of dish + potassium chlorate = 29·22 gm.
∴ Wt. of dish − potassium chlorate = 1·05 gm.
Wt. of water = wt. of solution − wt. of solid = 9·03 − 1·05 = 7·98 gm.

$$\therefore \text{Wt. of potassium chlorate in 100 gms. of water} = \frac{1{\cdot}05 \times 100}{7{\cdot}98} = 13{\cdot}2 \text{ gm.}$$

Temperature = 55° C.

THREE SOLUBILITY CURVES

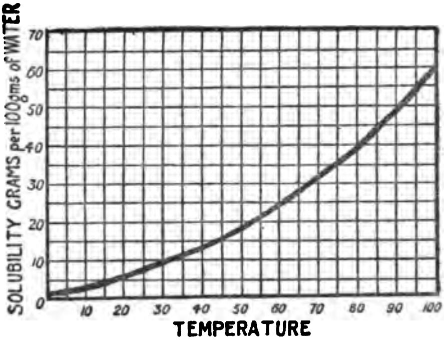

FIG. 29. POTASSIUM CHLORATE

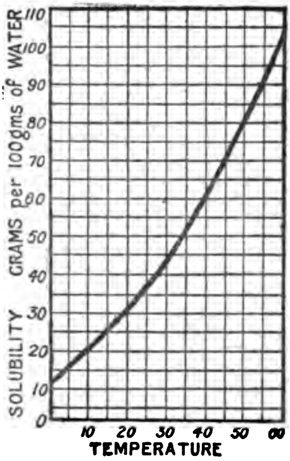

FIG. 30. NITRE

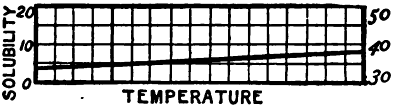

FIG. 31. COMMON SALT

EXERCISE

Make a solubility curve for alum.

Other Solvents.—Many bodies which are insoluble in water will yet dissolve in other liquids. **Alcohol**, *e.g.*, is a very useful solvent for such organic bodies as *quinine*. *Iodine* dissolves much more readily in alcohol than in water. Tincture of iodine is the alcoholic solution. **Ether**, again, will dissolve waxes and fats, which are insoluble in water.

EXERCISE

Devise a method for showing that when a slightly soluble substance is boiled with water some is dissolved.

e.g. Try if calcium sulphate (plaster of Paris) and lime dissolve to a small extent in water.

CHAPTER XIII

CRYSTALLISATION

MANY solutions, on cooling, deposit crystals of the dissolved substances. A hot solution, as we have seen, can usually contain more of the soluble substance than a cold solution, hence some of the dissolved matter must come out of solution on cooling, and the separated particles frequently collect together in **definite geometrical shape, viz. crystals.**

The same result may be obtained by allowing a solution to slowly evaporate. Large crystals can be obtained by slow cooling or evaporation; small ones result from more rapid treatment.

Exp. 52.—To obtain crystals of copper sulphate.

Dissolve some **blue vitriol** in water in an evaporating dish. Filter, if necessary, to clear the solution. Evaporate gently over a small flame. From time to time take out a few drops of the solution in a right-angled bent tube and cool under the water tap. While this is going on stop the heating of the dish. When a drop thus tested begins to form crystals in the bend, allow the contents of the dish to cool.

A sample of its contents has crystallised on cooling, and therefore we may expect the whole solution to do likewise.

When crystals appear pour off the liquid—called the **mother liquor**—drain the crystals and dry them on filter paper. Place the dry crystals in a watch-glass. Sketch several of the crystals. Notice that there is a distinct resemblance in each to a common type, and that there is some uniformity of shape to be observed.

EXERCISE

In the same manner obtain crystals of nitre, potassium chlorate, sodium sulphate, alum.

Sketch the crystals thus obtained.

By means of a protractor measure the angles between similar faces of several copper sulphate crystals. Do this for any other crystals you obtain which are large enough for measurement.[1] Tabulate your results.

Pick out good specimens of the crystals you have obtained. By means of a tiny fragment of sealing-wax and a long hair suspend them in beakers containing the mother liquors from which they are obtained. Place the beakers in a cool place and cover them with paper. Examine the crystals day by day. Brush off any small adhering crystals. Thus try to grow big crystals.

Examine the photographs of the crystals in Figs. 32–35, and if possible obtain actual specimens.

Fig. 32 illustrates copper sulphate, crystals of which you obtained in Exp. 51.

Fig. 33 shows the form taken by potassium ferrocyanide, a substance commonly called yellow prussiate of potash. In Fig. 34 we see octahedral (eight-sided) crystals of alum and prismatic crystals of nitre.

The specimens from which these photographs were taken were obtained on the manufacturing scale, by the slow evaporation and cooling of large masses of solution.

[1] Crystals of Iceland spar can easily be obtained for this purpose.

p. 144

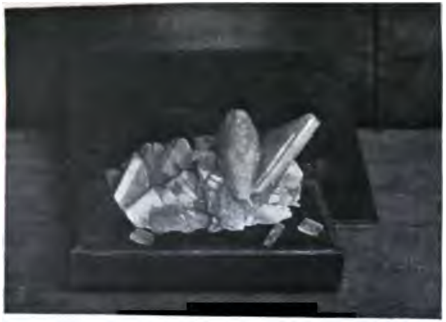

FIG. 32. COPPER SULPHATE

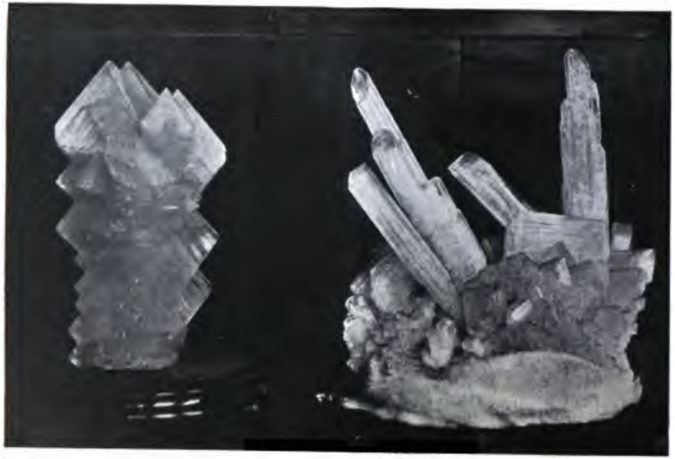

FIG. 34. ALUM AND NITRE

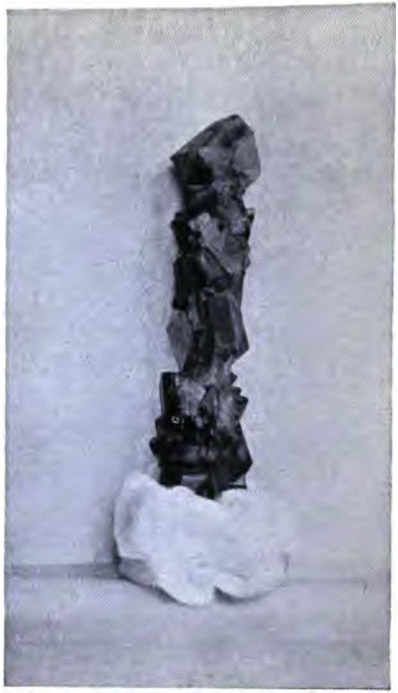

FIG. 33. POTASSIUM FERROCYANIDE

In Fig. 35 are shown specimens of rock salt, selenite, quartz, and Iceland spar.

Crystallisation is a very important means of obtaining pure substances. By repeated recrystallisation a very pure product is yielded. In this way the nitre for use in gunpowder is prepared, stirring and rapid cooling being used to obtain very small crystals, which are less liable to contain impurities than large ones.

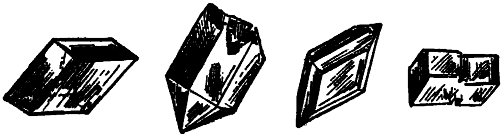

FIG. 35

In the previous instances we have seen that crystallisation takes place spontaneously on cooling a strong solution. If, however, special precautions are taken to prevent any disturbance of the solution, or any introduction of dust, etc., crystallisation may be deferred for some time.

This property is illustrated very well in the case of sodium thiosulphate.[1] If some of this salt be introduced into a flask and gently warmed, it melts in its *water of crystallisation* (see Exp. 52). Allow the solution to boil, and then plug the mouth of the flask with cotton-wool. Let the flask stand till cool, carefully avoiding shaking; then take out the plug and drop in a tiny fragment of the crystalline salt. The whole mass rapidly crystallises out, and it will be found that a considerable amount of heat is generated.

Exp. 52.—Water of crystallisation.

Heat a little powdered **alum** crystals in a dry bulb-tube. Notice that much water is evolved. This water is called the *water of crystallisation.* Hold a crystal in the flame. Notice that it seems to bubble and that steam comes off, leaving an amorphous mass. The loss of water is also accompanied by loss of crystalline shape.

Most crystals possess this water of crystallisation, and it may usually be easily driven off by heat. In some cases it is obstinately retained, however, even at high temperature. Common salt crystals have no water of crystallisation.

The dry substance thus left behind after heating is said to be **anhydrous.** We have already prepared anhydrous blue vitriol, and used it as a test for water.

The presence of water in blue vitriol may also be shown by simply immersing a crystal in oil of vitriol. This body has a great attraction for water, and in a short time the crystal is found to have turned white and anhydrous.

EXERCISE

Write on paper, using a solution of **cobalt nitrate.** This is a pink salt containing water of crystallisation. Gently warm the almost invisible writing over a bunsen flame. Notice that, as the water of crystallisation is expelled by the heat, a blue anhydrous substance remains, and the writing is rendered quite visible.

---

[1] The photographer's "hypo."

Some crystals readily lose their water of crystallisation, even if only exposed to the air. Such substances are called **efflorescent**. Notice that the large transparent crystals of washing soda rapidly lose their clear appearance, and become covered with a white crust. They have *effloresced.*

On the other hand, some crystalline substances absorb water with such eagerness that, if exposed to the air, they deprive it of moisture, and in time dissolve in this water they take up. **Copper nitrate** and **calcium chloride**, for example, soon become wet if exposed to the air. They are said to be **deliquescent**. We have already noticed the deliquescence of caustic soda (Exp. 33).

Deliquescent bodies are useful as *drying agents.* We have used anhydrous calcium chloride as a drying agent for hydrogen.

Exp. 53.—To find the percentage of water of crystallisation in barium chloride.

Weigh out a clean dry porcelain crucible and lid. Place in it about one gram of finely powdered crystals of barium chloride and weigh again. Place the crucible and lid on a pipe-clay triangle and heat strongly for fifteen minutes (Fig. 1).

Allow the crucible and contents to cool, and again weigh.

Repeat the heating for about five minutes, again cool and weigh.

If this last weighing agrees with the preceding one, we may take it that all the water has been evolved. If not, again heat for five minutes, and repeat if necessary till two successive weighings have the same value.

Enter your results as follows :—

Wt. of crucible and lid =20·65 gm.
Wt. of crucible and lid + barium chloride =21·65 gm.
∴ Wt. of barium chloride = 1·0 gm.
Wt. of crucible and lid + anhydrous barium chloride =21·50 gm.
∴ Wt. of anhydrous barium chloride = ·85 gm.
∴ Loss of water of crystallisation = ·15 gm.
∴ 1 gm. barium chloride contains ·15 gm. water of crystallisation ;
∴ 100 gm. ,, ,, ,, $\frac{\cdot 15 \times 100}{1} = 15$ gm. water of crystallisation.

EXERCISE

In a similar way find the percentage of water of crystallisation of alum and copper sulphate (avoiding decomposition which is seen when the substance begins to become brown in colour).

Exp. 54.—To find the percentage of water of crystallisation evolved by copper sulphate when heated to 100° C.

A convenient way to obtain a steady temperature of 100° C. is to use the steam from boiling water.

Weigh a dry clean watch-glass and place on it about one gram of finely powdered copper sulphate crystals (which have been purified by recrystallisation.) Weigh again. Support the watch-glass on the top of a small flask, half full of water. Place a wedge of paper between the flask and watch-glass, to allow free evolution of steam. Put a few pieces of paper into the water to help the boiling. Support the flask on wire gauze on a retort stand, and heat. The steam raises the temperature of the copper sulphate to 100° C.

After twenty minutes' heating remove the watch-glass, wipe off any adhering water, cool and weigh. Repeat the heating and weighing till the weight of the residue is constant.

Enter your weighings as follows:—

Wt. of watch-glass =4·98 gm.
Wt. of watch-glass + copper sulphate =5·9 gm.
∴ Wt. of copper sulphate = ·92 gm.
Wt. of watch-glass + anhydrous copper sulphate = 5·73 gm.
∴ Wt. of anhydrous copper sulphate = ·65 gm.
∴ Loss of water of crystallisation = ·27 gm.
∴ ·92 gm. copper sulphate contain ·27 gm. water of crystallisation;
∴ 100 ,, ,, ,, $\frac{\cdot27 \times 100}{\cdot92} = 29\cdot3$ gm. water of crystallisation.

EXERCISE

In a similar manner find the amount of water of crystallisation evolved at 100° C. from zinc sulphate (white vitriol).

CHAPTER XIV

ELEMENTS, MIXTURES AND COMPOUNDS

So far we have dealt with several substances which have undergone various changes, such as air, water, oxygen, hydrogen, copper oxide, etc.

Are these bodies similar in their composition? When oxygen and nitrogen are mixed together in the proper proportion, a gas is formed identical with ordinary air; but when two volumes of hydrogen and one volume of oxygen are exploded together we obtain water. Are air and water identical in their method of formation?

In the case of water we have a substance formed as a result of a **chemical action**, heat and light being evolved, and a quite different body formed, having *no properties in common* with those of the constituent gases.

In the case of air **no very striking action goes on.** Neither heat nor light is evolved, and the properties of the resultant substance are a kind of *mean or average* of those of its components. It supports combustion, but not so well as oxygen, nor, of course, is it a non-supporter of combustion like nitrogen. It is colourless, tasteless, odourless and neutral, as are both oxygen and nitrogen.

These two examples seem to suggest **two great classes** into which bodies may be placed, viz. **chemical compounds**, such as water and copper oxide, and **physical assemblies or mixtures**, such as air and gunpowder, in which we have the various constituents existing side by side, each adding its own properties to the sum-total.

Notice, however, that by suitable means both air and water can be separated into their components.

But there is a further group of substances, which, up to the present and with the means at our disposal, *have not been decomposed* into any simpler bodies.

Such bodies are termed chemical elements.

At different times many substances have been regarded as elements which upon further examination have been found to be capable of decomposition into two or more simpler constituents. Up to the time of **Davy** potash was thought to be an element. Davy, however, showed that it contained potassium combined with oxygen and hydrogen.

There are at present about *eighty* such elementary bodies, and others are continually being discovered.

Some of the common elements are:—

Non-Metallic: Oxygen, hydrogen, nitrogen, carbon, sulphur, phosphorus, etc.

Metallic: Copper, tin, zinc, iron, silver, mercury, potassium, sodium, etc.

A chemical compound is formed by the union of two or more elements, *e.g.* hydrogen and oxygen unite to form water, magnesium and oxygen form magnesium oxide.

Methods of Separation.—We have learnt that a chemical compound is usually very different from its constituent elements, and is separated with difficulty into them.

Strong heating, *e.g.*, is necessary to decompose mercury oxide into its component elements, oxygen and mercury.

But we should expect that a mixture would be separable without much difficulty, since its constituents exist side by side.

We have mentioned already some methods of separating mixtures into their parts, *e.g.* a mixture of a liquid with an insoluble solid can be **filtered**, as in Exp. 27.

Or if the insoluble solid is very heavy and sinks to the bottom of the vessel, then the clear *supernatant* (= swimming over or above) liquid can be poured off or **decanted.**

A mixture of two liquids, such as alcohol and water, can be **distilled**, the more volatile constituent (in this case alcohol) coming over first in the distillate, and the less volatile, viz. water, afterwards.

In a similar way a mixture of a volatile and a non-volatile solid may be broken up into its parts by heating, when the volatile substance is driven off, leaving the non-volatile body behind. This process is known as **volatilisation.**

A mixture of a soluble and an insoluble substance, again, may be parted by using a suitable solvent and filtering. The insoluble constituent forms the residue, and the soluble one is found in the filtrate, from which it may be obtained on subsequent **evaporation.**

Advantage may be taken also of such properties as density in separating mixtures. A very heavy component may be washed away from a less dense substance. In this way the mixture of sand and gold found in the beds of rivers is separated.

Exp. 55.—To separate a mixture of sand and salt.

Place the mixture in a test-tube. Add water and heat to boiling over a bunsen flame. Allow to settle and carefully **decant** into an evaporating basin, which is then placed on a tripod stand. **Evaporate to dryness** and obtain the salt. Pour water on the residue in the test-tube, boil, and again decant. Repeat this several times to wash the sand quite free from salt. Shake out the clean sand on to filter paper and dry it over a small flame.

This mixture has been separated and its constituents obtained by means of **solution, decantation** and **evaporation.**

EXERCISE

To find the percentage of the insoluble constituent in a mixture of a soluble and an insoluble substance (sand and nitre).

Weigh a clean, dry evaporating basin. Place about 2 gm. mixture in it, and weigh again. Add water, boil, decant off the clear solution. Repeat several times till the sand is clean. Dry the sand very carefully with a small flame. When cool, weigh. Calculate the percentage of sand in the mixture.

Exp. 56.—To obtain specimens of the constituents of gunpowder.

Place the gunpowder in an evaporating dish and add water. Support the dish on a tripod stand and heat to boiling. Pour the contents of the dish on to a filter paper in a funnel. Catch the filtrate in another dish. Evaporate the filtrate, gently testing the while for crystals with a right-angled bend. When crystals appear in the bend allow the dish and contents to cool. Pour off the mother liquor and dry the crystals of **nitre.**

Pour boiling water on to the residue in the funnel. Repeat this washing several times till all the nitre is washed out. When the residue is thus obtained quite free from nitre, take out the paper, support it on a wire gauze on the retort stand, and place a small flame below to dry it.

We have now a dry mixture of **sulphur and charcoal.**

To separate these we take advantage of the fact that *sulphur is volatile*, while charcoal is not. Place some of the residue in a bulb-tube and heat. The sulphur melts and finally vaporises, condensing again as a *sublimate* on the sides of the tube. When all the sulphur has volatilised, allow the tube to cool, and by means of a file-scratch break the tube between the bulb and the sublimed sulphur.

We have thus obtained specimens of nitre, sulphur, and charcoal, the constituents of gunpowder, by the processes of **solution, filtration, evaporation** and **volatilisation.**

EXERCISE

To find the percentage of nitre in gunpowder.

Weigh out about 2 gm. gunpowder. Place it in a test-tube half full of water. Boil and filter into a clean, dry, weighed evaporating dish.

Wash the residue on the filter paper with a little boiling water, allowing the washing water to filter into the dish. Repeat the washing with successive small amounts of boiling water.

Then evaporate the contents of the dish to dryness. Avoid loss by spurting. When all the water has been driven off allow the dish to cool, and weigh again.

Thus find the percentage of nitre in the gunpowder.

An instructive illustration of the fact that mixtures are simply *physical assemblies* in which the components may be regarded as lying side by side is to be found in the process of separation of a mixture of two soluble salts by *fractional crystallisation*. This method depends on the superior solubility o one constituent.

Make up a mixture of potassium chlorate and potassium dichromate. Notice that the resulting substance has the average colour of the two salts.

Dissolve in the least amount of boiling water and allow to stand till cool.

Part of the less soluble potassium chlorate crystallises out, but with it always appears some of the more soluble dichromate. Decant off the mother liquors and drain the crystals, which are reddish in shade, and again dissolve in the least amount of boiling water. This second crystallisation affords potassium chlorate, which after draining is only feebly tinged with red dichromate. A third crystallisation gives the pure salt.

EXERCISES

1. Describe how to get a specimen of alcohol from wine.
2. How would you separate a solution of saltpetre in dilute nitric acid so as to obtain nitric acid free from saltpetre?
3. How would you free sulphuric acid from dissolved hydrochloric acid?

CHAPTER XV

THE QUANTITATIVE COMPOSITION OF WATER

We found that water contained both oxygen and hydrogen, but we do not yet know *in what proportions* these elements are combined. We may find the composition of water by **volume** or by **weight**, *i.e.* the proportions by volume or by weight in which the two constituents are combined. Again, each method of stating this composition may be determined by **analysis** or **synthesis**.

1. Composition of Water by Volume.

A. Analytically.—When a current of electricity is sent through slightly acidified water by means of platinum wires connected to a battery, bubbles of gas are seen to be evolved from the ends of the wires or **electrodes.**

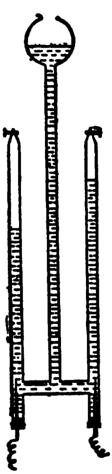

FIG. 36

Perfectly pure water is a non-conductor of electricity, so a little acid is added to allow the current to pass.

The evolved gases may be collected by holding test-tubes full of water over the electrodes. A more convenient apparatus is seen in Fig. 36.

The electrodes here are platinum plates connected to platinum wires, which pass into two vertical tubes provided with taps. These tubes serve to collect the gases. They are joined near their ends to a *third vertical tube*, ending in a bulb. Before the current is passed, the taps are opened, and acidified water is poured into the bulb till both tubes are filled. The taps are then closed, and the electrodes are connected with the battery. As the gases given off from the platinum plates accumulate, the displaced water is driven up into the bulb.

One tube is seen to be filling much more rapidly than the other; and when it is full, the other is seen to contain only half as much gas. Disconnect the battery, and hold a *lighted taper* over each tap as it is opened. The pressure of the water in the bulb forces the gas out, and it is noticed that the gas which accumulates in the larger amount catches fire and burns with a blue flame. It is evidently **hydrogen.**

The other gas supports combustion, and is **oxygen.**

This experiment shows that when water is analysed by electricity or **electrolysed**[1] there are evolved **two volumes of hydrogen for each volume of oxygen.**

B. Synthetically.—**Cavendish,** in 1781, found that when air and hydrogen were exploded together in various proportions water was formed and some of each gas disappeared. When the gases were in the proportions of 423 volumes hydrogen and 1,000 volumes air, then nearly all the hydrogen and one-fifth of the air vanished, leaving a dew of water behind.

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[1] The process is termed Electrolysis.

Now he had previously shown that air contained about one-fifth of its volume of oxygen, so that 423 volumes hydrogen had evidently combined with about 200 volumes oxygen.

On repeating this experiment, but using oxygen instead of air, he found that when 2 volumes hydrogen were exploded with 1 volume oxygen all the gas vanished, leaving water.

Exp. 57.—To show that 2 volumes hydrogen combine with 1 volume oxygen to form water.

Take a *eudiometer* (Fig. 37) graduated in c.c. along the closed limb, and having two platinum wires sealed through the glass near the end.

Pour mercury in till the closed limb is full and mercury stands in the bend up to the tap.

Electrolyse some acidulated water (p. 153), and connect the two taps of the apparatus by rubber tubing with two ends of a glass T-piece, so that a stream of the mixed gases may be led from the third end. Connect this end with a delivery tube, which can dip right under the mercury in the eudiometer. Place a few inches of rubber tubing on the end of the delivery tube to allow the gases to be delivered into the closed limb.

When all is ready, open the taps and allow half the gas to escape rapidly, thus driving out the air from the tubes. Then collect about 20 c.c. in the eudiometer. Run out some mercury from the open limb so as to lower the pressure.

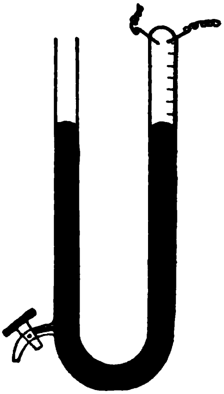

FIG. 37

Hold your thumb over the open end, and **pass a spark through the gas** by means of an induction coil. An explosion takes place, accompanied by a bright flash. On removing the thumb the mercury rushes up into the closed limb and fills it. The water formed is such a small amount as to be practically invisible. Evidently, then, **2 volumes hydrogen and 1 volume oxygen, on explosion, disappear,** a very small volume of water being formed.

To find what proportion of water is formed we must get it in the gaseous condition, *i.e.* the eudiometer must be kept at a *temperature over* 100° C.

Exp. 58.—To show that 2 volumes hydrogen and 1 volume oxygen combine to form 2 volumes steam.

Take a eudiometer, as in Exp. 105, and place in it about 20 c.c. of the mixed gases obtained by electrolysing water. The closed limb of the eudiometer is provided with a jacket, *i e.* a wider tube fitted to it by means of corks which carry right-angled bends (Fig. 38).

Through the upper bend enters the vapour of **amyl alcohol,** which is boiling in a flask connected with the bend. This liquid boils at 132° C., thus the gases in the eudiometer are kept at a temperature over 100° C. The condensed amyl alcohol escapes from the bend at the lower end of the jacket.

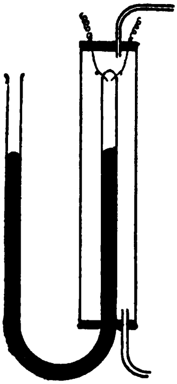

FIG. 38

When the eudiometer has arrived at a steady temperature pour mercury into the open limb till it stands at the same level in each limb. Measure off the volume of the mixed gases, which are thus at **atmospheric pressure.** Lower the pressure as before.

Hold the thumb over the open limb **and pass a spark through the gas.**

Combination takes place with a bright flash. Remove the thumb and pour mercury into the open limb till the levels are again coincident.

Read off the volume of the remaining steam. It is found to be two-thirds of the volume occupied by the hydrogen and oxygen.

Allow the eudiometer to cool, and notice that the mercury rises in the closed limb as the steam condenses, finally filling the whole tube.

Thus **2 volumes hydrogen and 1 volume oxygen unite to form 2 volumes of steam.**

2. Composition of Water by Weight.

A. Analytically.—This experiment was first performed by **Lavoisier**, who passed steam over iron contained in a tube heated red hot in a furnace (Exp. 43, Fig. 22). The steam was decomposed, forming black oxide of iron and liberating hydrogen. He condensed any steam which escaped analysis and found its weight. He weighed the water which he turned into steam to pass over the iron. Thus he obtained the actual weight of the water which was decomposed. By weighing the tube containing the iron before and after the experiment he obtained the increase in weight. This increase represents the weight of **oxygen** in the water.

Thus he obtained the weight of oxygen contained in 100 parts of water, and by subtraction the percentage of **hydrogen.**

The method is inaccurate, because some of the water escapes condensation.

B. Synthetically.

Exp. 59. —Fit up a flask for the evolution of hydrogen, putting into it pure zinc foil and a few drops of copper sulphate solution, which promotes the liberation of the gas. By means of right-angled bends connect the flask with a smaller one, A, containing **strong sulphuric acid.** This serves to dry the hydrogen as it bubbles through. Connect the drying flask with a reduction bulb, B, containing some freshly heated **copper oxide.** The water formed in the reaction is absorbed in a U-tube C, containing **calcium chloride**, which is attached to the reduction bulb.

Weigh the reduction bulb and the U-tube. Pour dilute sulphuric acid down the thistle funnel. Allow the gas to sweep out all the air from the apparatus, testing it from time to time. When pure hydrogen issues from the end of the U-tube, *heat the copper oxide gently*. Water is formed and condenses on the cold sides of the bulb-tube. As the temperature increases it is driven into the U-tube and there absorbed. When all the copper oxide is reduced, stop heating but allow hydrogen to pass *till the apparatus is cold*. This ensures that all the water formed has been carried into the U-tube.

Disconnect the apparatus, and weigh the bulb-tube and U-tube again.

The decrease in weight of the reduction bulb is due to **loss of oxygen**; the increase in weight of the U-tube is due to the **water** formed.

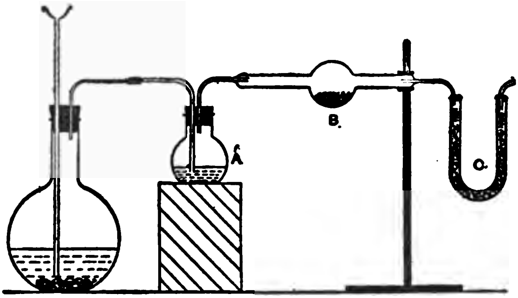

FIG. 39

Thus the percentage of oxygen in water can be found, and by subtraction the percentage of hydrogen.

Enter the results as follows:—

Wt. of reduction bulb and copper oxide	= 30·56 gm.
Wt. of reduction bulb + copper	= 29·5 gm.
∴ Wt. of oxygen	= 1·06 gm.
Wt. of calcium chloride tube before experiment	= 12·03 gm.
Wt. of calcium chloride tube after experiment	= 13·28 gm.
∴ Wt. of water	= 1·25 gm.

∴ 1·25 gm. water contain 1·06 gm. oxygen.

$$\therefore \text{100 gm. water contain } \frac{1{\cdot}06}{1{\cdot}25} \times 100 = 84{\cdot}8 \text{ gm. oxygen.}$$

∴ Percentage of hydrogen = 15·2.

Accurate experiments give the percentage of oxygen as 88·8.

CHAPTER XVI

MEASUREMENT OF THE VOLUME OF GAS EVOLVED IN A REACTION

THE simplest method of measuring the volume of a gas evolved in a reaction would be to collect the gas over water in a graduated jar, supposing the gas to be insoluble in water.

But there are several drawbacks to this, *e.g.* some gas might be lost during the introduction of the materials needed for its evolution, or if the volume evolved was only small, a graduated jar would not be a suitable vessel in which to collect it.

Exp. 60.—To find the amount of gas evolved when 1 gm. zinc is dissolved in dilute hydrochloric acid.

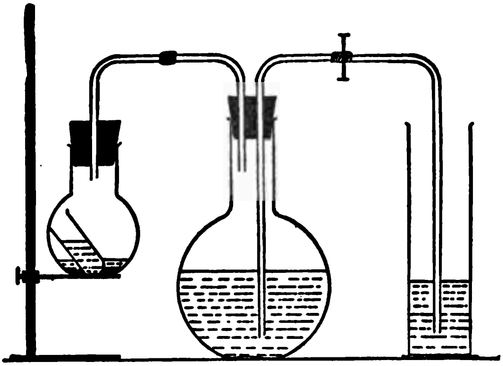

FIG. 40

Take apparatus as shown in Fig. 40 above. A small flask is provided with a well-fitting cork through which passes a right-angled bend. This is connected by rubber tubing with a similar bend passing through the cork of the large flask. Through this cork also passes a long right-angled bend, dipping almost to the bottom of the flask, and connected by rubber tubing with a similar long bend which passes into a gas jar. A

clip is fastened to the rubber, between these two bends, which form a siphon.

The large flask is filled with water.

A piece of zinc foil is weighed out. Its weight should not exceed ·5 gm. It is placed in the small flask and covered with water. A little test-tube, made by drawing off the top part of a test-tube in the blowpipe flame, is nearly filled with strong **hydrochloric acid.** By means of a piece of thread it is suspended from the cork of the small flask.

We have, then, in the small flask everything ready for the evolution of hydrogen, without any chance of loss if the apparatus is fitted carefully. Open the clip and blow through the short bend of the large flask. This fills the siphon tube with water. Close the clip. Connect up the apparatus as in Fig. 40. Then the gas, when evolved, will drive **its own volume of water** from the flask into the gas jar, so that we really shall measure a volume of water.

But it is very important that after the experiment *the conditions of temperature and pressure shall be the same* as they are to begin with, because a gas volume is very easily altered by a change of temperature or pressure. So we must not measure the volume of water till the apparatus is cool, and also we must see that the pressure in the apparatus is the same before and afterwards. It is easy to arrange that this pressure shall be atmospheric, in the following way.

Pour some water in the gas jar. Open the clip and raise or lower the jar till the levels of the water in the flask and jar are equal. Close the clip. The pressure in the apparatus is now atmospheric. Empty the water out of the jar. Open the clip, shake the small flask till the acid is spilt on to the zinc. Hydrogen is evolved, and it displaces its own volume of water from the big flask into the jar. When the zinc is dissolved, allow the apparatus to cool down. Lower the gas jar till the levels of the water in it and in the flask are again the same. Close the clip. The pressure inside the apparatus is again atmospheric.

Pour the water in the jar into a measuring cylinder, and find its volume.

Find the atmospheric pressure and the temperature of the room.

Correct the volume you have found to N.T.P. (normal temperature and pressure, 0° C. and 760 mm. of mercury).

EXAMPLE

·35 gm. of zinc evolved 123 c.c. gas.

Barometer reading 756 mm. Temperature 11·5° C.

∴ Absolute temperature = 284·5°.

$$\therefore 1 \text{ gm. zinc evolves } \frac{123 \times 1}{\cdot 35} \text{ c.c.} = 350 \text{ c.c.}$$

$$350 \text{ c.c. at } 756 \text{ mm.} = 350 \times \frac{756}{760} \text{ c.c. at normal pressure.}$$

$$\frac{350 \times 756}{760} \text{ c.c. at } 11{\cdot}5^\circ \text{ C.} = \frac{350 \times 756}{760} \times \frac{273}{284{\cdot}5} \text{ c.c. at N.T.P.}$$

$$= 334 \text{ c.c.}$$

EXERCISE.

Weigh out about ·2 gm. aluminium wire, and find the volume of hydrogen evolved when it is dissolved in dilute hydrochloric acid.[1]

Hence calculate the amount given off when 1 gm. dissolves at N.T.P.

Similarly find the volume of gas evolved from 1 gm. iron filings when dissolved in dilute sulphuric acid.

Exp. 61.—To find the volume of gas evolved when 1 gm. marble dissolves in dilute hydrochloric acid.

In this case carbonic acid gas will be given off. Now this gas is soluble in water to a considerable extent, so the apparatus used above must be modified. One method is to place a layer of **paraffin oil** on the surface of the water in the large flask. This prevents the gas from coming into contact with the water, and it is insoluble in the oil. A better way is to **saturate the water in the flask with carbonic acid gas.** Place in the water a few crystals of washing soda and add a little hydrochloric acid. Carbonic acid gas is evolved, and bubbles through the water making a saturated solution. When the bubbling ceases connect up as before, having placed 1 gm. of marble in the small flask. Level off as before, unclip and upset the acid. When the action is over, allow to cool; level off again, and measure the volume of the displaced water. Correct for temperature and pressure, and calculate the volume of gas evolved when 1 gm. of marble is dissolved in dilute hydrochloric acid.

EXERCISE

In a similar way find the volume of carbonic acid at N.T.P. evolved from 1 gm. of whitening and from 1 gm. washing soda.

Exp. 62.—To find the volume of gas evolved when 1 gm. magnesium dissolves in dilute hydrochloric acid.

A small mass of magnesium evolves a large volume of gas when dissolved in dilute acid. Consequently the apparatus used for zinc and marble will not be suitable unless the weighing of the magnesium is made with very great care.

It is more convenient to find the weight of, say, 1 metre of magnesium ribbon and then to cut off a certain length, say, 10 cm., and to use the apparatus shown in Fig. 41.

Take a **burette** and fix a funnel to the tap by means of rubber tubing. Fill the burette with water and invert it in a sink containing water, so that the end of the burette rests on the sink. Support the funnel in a ring of a retort stand. Bend the 10 cm. of magnesium ribbon and place it in the burette near the end. The elasticity of the ribbon holds it in place.

Open the tap so that the water in the burette falls to the first mark on the graduations (usually indicating 50 c.c.). Pour strong hydrochloric acid into the funnel and allow it to enter the burette slowly through the tap, taking care that no air is allowed to get in. As the acid falls through the water it comes to the magnesium. Hydrogen is rapidly evolved and collects in the burette. When the apparatus is cool pour water into the sink till the level is the same inside the burette and out. Read off the graduation at which the water stands. Suppose this is 10 c.c. Then 40 c.c. of hydrogen have been evolved.

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[1] Heat may be required to start the action.

Correct this volume for temperature and pressure, and calculate what volume at N.T.P. would be given off when 1 gm. magnesium dissolved in dilute acid.

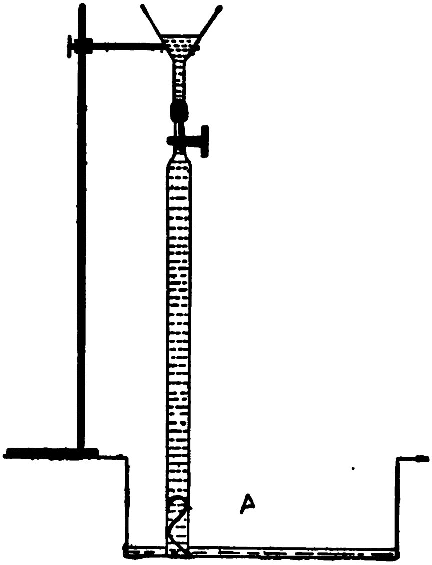

FIG. 41

EXERCISE

Instead of the small flask in Exp. 60, fit on to the large flask a hard glass tube which has been previously weighed. Weigh into it some powdered potassium chlorate (not more than ·3 gm.).

Heat gently at first and more strongly afterwards, thus driving off oxygen. Allow to cool and measure the volume of water displaced. Thus find the volume of oxygen given off when 1 gm. potassium chlorate is heated.

Weigh the tube again when cool. This gives the weight of oxygen. Then find the density of oxygen.

Given that 1 gm. marble gives on heating ·44 gm. carbonic acid gas, find the density of this gas.

CHAPTER XVII

FURTHER EXAMINATION OF COMMON SUBSTANCES

Sulphur.—This element is usually seen either as a yellow powder, viz. **flowers of sulphur**, or as brittle sticks or lumps, viz. **brimstone.**

Examine a little sulphur. Notice that brimstone, on being broken, shows distinct signs of a **crystalline** structure. A stick of brimstone held in the hand usually breaks into pieces owing to its being a very **bad conductor** of heat. Brimstone means *burning stone.*

Expt. 62.—Effect of heat on sulphur.

Heat some flowers of sulphur in a dry test-tube. Hold the tube at first *well above the bunsen flame.* Notice that the sulphur gradually melts to a **pale amber coloured liquid,** which is **mobile,** *i.e.* easily flows about.

Heat a little more strongly. The sulphur now **darkens in colour** and loses its mobility. As the temperature rises it becomes almost black, and then is extremely **viscous** or sticky. Invert the tube, and notice that the liquid does not fall out. All the while vapour is being given off, which condenses on the cold sides of the tube as a yellow powder. This yellow powder, when produced similarly on a large scale, is called **flowers of sulphur.**

Presently the sulphur boils. Observe that the boiling liquid is more mobile. When all the sulphur has boiled away notice the **deep brown coloured vapour** in the tube. As the tube cools these changes present themselves in the *reverse order*, first a black mobile and then a viscous liquid, followed by a yellow mobile fluid which soon solidifies to a yellow mass.

Sulphur readily combines with copper and with many other metals, forming **sulphides.** It burns in air or oxygen, and yields an acid oxide.

Sulphur is insoluble in water, but easily dissolves in carbon disulphide.

Expt. 63.—Pour a few c.c. of carbon disulphide into a test-tube. Remember that this liquid is extremely inflammable. Add a little sulphur. When dissolved pour the solution into a watch-glass and allow it to evaporate slowly. Curious tree-like growths frequently appear at the edge of the glass, and these, if examined with a lens, are seen to be composed of crystals of sulphur.

Iron Pyrites.—This is a common, widely diffused *mineral*, which is often found associated with coal and chalk. It has a **brassy, metallic lustre**, and is very **dense**. On exposure to air and moisture iron pyrites gradually **oxidises**, eventually forming *green vitriol*.

Examine a lump of iron pyrites. Observe its **crystalline** nature, and notice that the crystals are *cubical*. Heat some in a bulb-tube. There is a crackling, or **decrepitation**, due to the breaking of the crystals by unequal expansion. It is a **bad conductor** of heat. Sulphur is deposited on the walls of the tube and a gas with a choking smell is evolved.

Heat a little pyrites in a piece of glass tubing open at each end and held in a sloping position. Notice red oxide of iron left and the smell of burning sulphur.

Iron pyrites is not used as a source of iron, but its sulphur is utilised in the manufacture of sulphuric acid. The pyrites is roasted in air, the sulphur is converted into its oxide, and this is further oxidised and combined with water to form the acid.

Green Vitriol, Iron Sulphate.—This salt can be obtained by the slow oxidation of iron pyrites. It crystallises in large, glassy crystals, hence the name vitriol. It can be used as a means of preparing sulphuric acid.

Heat a little powdered green vitriol in a bulb-tube. A large amount of *steam* is evolved, some of which condenses on the sides of the tube. White anhydrous amorphous iron sulphate remains. Heat this strongly. Acid fumes are given off, and red oxide of iron remains in the bulb.

Try to prepare green vitriol by dissolving iron filings in dilute sulphuric acid, filtering and evaporating.

Blue Vitriol, Copper Sulphate.—This substance, frequently called **blue stone,** is usually obtained in fine blue crystals of a rhombic shape (Fig. 32). It may be obtained in a similar manner to green vitriol from the oxidation of copper pyrites.

Heat some powdered blue vitriol in a bulb-tube. Notice again that white *amorphous, anhydrous copper sulphate* is left, which we have used as a test for water. Heat the bulb more strongly. Notice that the residue darkens, owing to the formation of **copper oxide.** Test with blue litmus paper, to see if any acid vapours are evolved.

For the estimation of the water of crystallisation of blue vitriol, see Expt. 53.

In the case of this substance the water of crystallisation is not all driven off at 100° C. The remaining water is not evolved till a high temperature is reached.

White Vitriol, Zinc Sulphate.—This salt may be obtained when zinc is dissolved in dilute sulphuric acid. Prepare a crystalline specimen of zinc sulphate in this way. Notice the **long needle-shaped crystals.** Heat some in a bulb-tube. Water of crystallisation is evolved, and eventually **zinc oxide** left.

Also prepare zinc sulphate, using zinc oxide and sulphuric acid.

Alum.—This common substance is a compound of **potassium** or **ammonium sulphate** and **aluminium sulphate,** and is used in large quantities in dyeing. Recrystallise some commercial alum. Pick out the best crystals and grow them. Notice their regular eight-sided shape. Such crystals are called **octahedral** (= eight-sided). Usually the sharp edges of the crystals are bevelled off by sets of parallel planes. Find the effect of heating a crystal of alum in a bulb-tube. Examine some **chrome alum.** Recrystallise this substance, and compare its crystalline form with that of ordinary alum.

Sal-ammoniac. Ammonium Chloride.—This body is usually met with in the form of a *white crystalline powder* or in hard fibrous lumps. It is *extremely soluble* in water.

Heat some in a bulb-tube. Notice that it is very volatile and condenses as a **sublimate** on the cool sides of the tube.

Heat some in a test-tube with lime. A gas is evolved having a very pungent smell. This is **ammonia gas.** Fix a cork provided with a long right-angled bend into the test-tube and allow the ammonia to pass into a little water contained in another tube. Notice that the solution smells of the gas and is very alkaline.

Washing Soda, Sodium Carbonate.—This salt was referred to by **Black** as *mild alkali.* It is very soluble in water, and like all carbonates, evolves carbonic acid gas when treated with acids. When boiled with lime it yields caustic soda and chalk. Examine some crystals of washing soda. Notice that some have lost their clearness and have become opaque, being covered with a white powdery substance; washing soda is **efflorescent.**

Heat some in a bulb-tube. Notice that a large amount of **water of crystallisation** is evolved, and white anhydrous, amorphous sodium carbonate remains.

Act on a little washing soda in a test-tube with dilute acid. Test the evolved gas with a drop of lime-water held on the end of a glass tube.

Nitre, Saltpetre, or Potassium Nitrate.—Nitre is found as a crust upon the earth in certain parts of South America and India. Recrystallise some commercial nitre, cool slowly so as to obtain large crystals. Notice the needle-like shape. Examine a crystal through a lens, and make a sketch.

Heat some powdered nitre in a bulb-tube. The crystals melt, but *no water of crystallisation* is evolved. An effervescence is noticed. Test the gas which is being evolved with a glowing splint. It bursts into flame because oxygen is given off.

Heat some nitre with a few drops of strong sulphuric acid. Nitric acid is formed. Add a little copper and notice the brown fumes evolved.

Recognition of Gases evolved in a Reaction. If a gas or vapour is evolved when a substance is heated or treated with an acid, apply the following tests:—

1. **Appearance.**

   (*a*) If the vapour condenses on the sides of the tube, it is probably *water*. Confirm by using anhydrous blue vitriol.

   Evolution of **brown fumes**, and also a peculiar streakiness on the tube, point to the liberation of *nitric acid*. Add a little copper, and notice if a dense red gas is evolved.

2. **Smell.**

   (*a*) **Acid fumes** may show *nitric acid* (see 1).

   (*b*) **A choking, irritating gas** may be *sulphur oxide*. Test with blue litmus paper.

   (*c*) **Ammonia** may be easily recognised by its peculiar smell. Confirm by (1) its alkaline action on litmus; (2) fuming with hydrochloric acid.

3. **Action on a lighted taper.**

   (*a*) If the gas burns, it is probably *hydrogen*. Hold a cold surface over the flame, and see if a dew is deposited.

   (*b*) If the taper burns much more brightly, *oxygen* may be suspected. Test with a glowing splint.

   (*c*) If the taper is extinguished, **carbonic acid gas** may be coming off. Test with a drop of lime-water suspended from a glass tube. Nitrogen also will not support combustion, but has no effect on lime-water.

APPENDIX

FITTING

To cut glass tubing.

Take a length of glass tubing (No. 4 or 5), and draw a triangular file smartly across it about 20 cm. from the end. Snap off the tube at this mark, **slightly pulling the tube** at the same time.

Hold the ends of the tube in turn in a bunsen flame, remembering that the tip of the inner cone is the hottest part. Slowly **revolve the tube** till the glass just melts at the edge. This destroys the sharp edge left after cutting. Glass tubing for apparatus should always have the ends rounded off in this fashion.

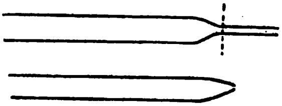

FIG. 42.

To make a jet.

Draw out a piece of glass tubing as before. Break off the capillary tube where it joins the wider tube, and heat the end in the flame, constantly turning the tube till the aperture is of the desired dimensions (Fig. 42).

To make a right-angled bend.

Take a piece of glass tube about 20 cm. long. Round off the ends. Hold it by one end lengthwise in an ordinary **bat's-wing flame**, keeping it constantly rotating.

After a time the glass softens and begins to bend downwards by its own weight. Allow this to go on till a right angle is formed. Take the tube out of the flame, and hold it up till the angle is seen against a window-frame. Move the free end about till an exact right angle is made, using the window-frame as a kind of **set-square**.

Take care not to displace one end of the tube out of plane with the rest of the tube while soft.

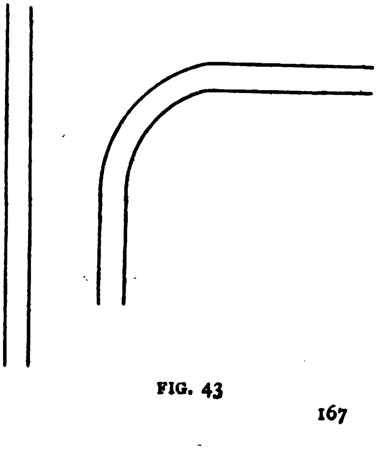

FIG. 43

Make a long right-angled bend by taking a piece of tubing about 30 cm. long and holding it in the flame 10 cm. from one end.

A bunsen flame is not used for bending glass, because it only heats a very short piece of tube. **We want a long piece heated, so as to form a *gradual, even curve.***

To bore a cork.

Take a cork a little too large for the vessel used. Roll it under the sole of your boot till it becomes *soft*. See that it will now fit the vessel. If not, squeeze it again.

Take a cork-borer, *i.e.* a brass tube, sharpened at one end and provided with holes near the other end, through which a rod is passed as a handle.

Moisten the borer, and begin to bore at the narrow end of the cork. Take care to keep the borer parallel with the long axis of the cork, and use a screwing motion in driving the borer. It is often useful to hold an old, used cork against the other end of the cork you are boring, otherwise, if the borer is a little blunt, a piece is frequently dragged out of the cork when the borer comes through. If two holes have to be bored, take care that each is well clear of the other. Begin boring *near the circumference of the narrow end.*

The holes can be enlarged, if necessary, by a rat's-tail file. While filing **roll the cork backwards and forwards**, so that the enlarging is uniform.

Use of rubber tubing.

Small pieces about 5 cm. long are cut off from a length of the tube. Rubber tubing can be more easily fitted to glass when it is moistened.

If it is too large for the tube, double it back on itself.

For fitting a glass tube into a reduction bulb a piece of **bunsen tubing** is slipped on to the end of the bulb-tube.

If the rubber be now doubled back on itself an ordinary piece of glass tube can be fitted into it.

Very often a reduction bulb or a U-tube can be fitted without corks by simply placing a piece of thick rubber tubing on the end of the right-angled bend (*e.g.*) to be attached to the tube, so that it fits tightly.

If rubber tubing become hard place it in hot water and rub it between the fingers.

Plenty of examples of fitting are to be found in the previous pages.

Precautions to be taken with apparatus.

Never heat flasks or retorts containing liquids with the naked flame. Use wire gauze to diffuse the heat.

Begin heating gently, move the flame about at first so that the vessel is warmed gradually.

Similarly allow flasks, etc., to cool gradually.

Do not put hot glass on cold surfaces.

Remember not to weigh vessels which are too hot to be easily handled.

Clean out test-tubes, etc., immediately after use.

TABLES OF PHYSICAL CONSTANTS

I. METRIC TABLE OF LENGTHS

10 millimetres = 1 centimetre.
10 centimetres = 1 decimetre.
10 decimetres = 1 METRE.
10 metres = 1 dekametre.
10 dekametres = 1 hectometre.
10 hectometres = 1 kilometre.

II. TABLE OF AREAS

100 sq. millimetres . . . = 1 sq. centimetre.
100 sq. centimetres . . . = 1 sq. decimetre.
100 sq. decimetres . . . = 1 sq. metre.
100 sq. metres . . . . = 1 sq. dekametre.
100 sq. dekametres . . . = 1 sq. hectometre.
100 sq. hectometres . . . = 1 sq. kilometre.

The are, which is 1 sq. dekametre, *i.e.* 100 sq. metres, is used for measuring land.

III. TABLE OF VOLUMES

1,000 cu. millimetres . . = 1 cu. centimetre.
1,000 cu. centimetres . . = 1 cu. decimetre [called a litre].
1,000 cu. decimetres . . = 1 cu. metre.

Another system of measuring volume is that in which the litre is taken as the unit.

10 centilitres = 1 decilitre.
10 decilitres = 1 litre.
10 litres = 1 dekalitre.
10 dekalitres = 1 hectolitre.
10 hectolitres = 1 kilolitre.

IV. TABLE OF MASSES

10 milligrams = 1 centigram.
10 centigrams = 1 decigram.
10 decigrams = 1 gram.
10 grams = 1 dekagram.
10 dekagrams = 1 hectogram.
10 hectograms = 1 kilogram.
1000 kilograms = 1 tonne.

V. CONVERSION TABLES

1 metre = 39·37 inches.
1 cm. = 0·3937 inch.
1 kilometre = 0·62 mile.
1 sq. cm. = 0·155 sq. in.
1 c.c. = 0·061 cu. in.
1 gm. = 15·432 grains.
1 kilogram = 2·205 lbs.
1 litre = 1·76 pint.

1 inch = 0·0254 metre.
1 inch = 2·54 cm.
1 mile = 1·609 kilom.
1 sq. in. = 6·45 sq. cm.
1 cu. in. = 16·386 c.c.
1 grain = 0·065 gm.
1 lb. = 453·593 gm.
1 pint = 568·23 c.c.

VI. USEFUL CONSTANTS

1 cubic foot of water at 39° F. or 4° C. weighs 62·41 lbs.

Pressure of 1 atmosphere = 14·7 lbs. per sq. in. = 2,116 lbs. per sq. ft. = 760 mm. of mercury.

Length of a seconds pendulum (in London) = 39·139 ins. or 99·41 cm.

$\sqrt{2} = 1{\cdot}41421\ldots$, $\sqrt{3} = 1{\cdot}73205\ldots$, $\sqrt{5} = 2{\cdot}23606\ldots$

A gallon of pure water at 62° F. weighs 10 lbs.

VII. SPECIFIC GRAVITIES

SOLIDS

Aluminium	2·7	Iron (wrought)	7·8
Brass	8·4	Lead	11·4
Bronze coinage	8·7	Marble	2·7
Coal	1·25	Slate	2·5
Copper	8·9	Sulphur (roll)	2·0
Cork	0·24	Tin	7·3
Glass (crown)	2·5	Wax (bees')	0·96
Glass (flint)	3·3	Wax (paraffin)	0·89
Ice	0·92	White sand	2·6
Iceland spar	2·7	Zinc	7·2
Iron (cast)	7·2		

LIQUIDS

Alcohol	0·79	Milk	1·03
Benzol	0·88	Olive oil	0·92
Glycerine	1·26	Petroleum	0·82
Mercury	13·596	Turpentine	0·87

VIII. MELTING POINTS

Beeswax	62° C.	Paraffin wax	52° C.
Butter	33° C.	Tin	233° C.
Lard	33·2° C.	Zinc	420° C.
Lead	325° C.		

IX. BOILING POINTS

Alcohol (absolute)	78·2° C.	Ether	35·5° C.
Alcohol (methyl)	66·3° C.	Glycerine	290° C.
Benzol	80·4° C.	Mercury	350° C.
Calcium chloride, sat. sol.	179·5° C.	Sodium chloride, sat. sol.	108·5° C.
Carbon bisulphide	48·0° C.	Turpentine	156° C.

X. SPECIFIC HEATS

LIQUIDS

Alcohol (ethyl)	0·615	Mercury	0·033
Alcohol (methyl)	0·613	Olive oil	0·47
Benzene	0·43	Turpentine	0·467
Glycerine	0·576		

SOLIDS

Aluminium	0·21	Lead	0·031
Brass	0·094	Marble (white)	0·214
Copper	0·095	Paraffin wax	0·67
Glass (crown)	0·16	Sulphur	0·184
Glass (flint)	0·12	Tin	0·056
Iron	0·113	Zinc	0·0935

XI. LATENT HEATS OF FUSION

Beeswax	42·3	Paraffin wax	35·00
Ice	79·25		

XII. LATENT HEATS OF VAPORISATION

Alcohol (ethyl)	209	Turpentine	69
Alcohol (methyl)	264	Water	536
Ether	90·5		

GLOSSARY AND DEFINITIONS

PHYSICS

Absolute Temperature. Temperature measured on the *absolute scale*, of which −273° C. is the zero.

Absolute or Real Expansion of a Liquid. It is the expansion which would be observed if the containing vessel did not expand. See *Apparent Expansion*.

Apparent Expansion of a Liquid. A liquid must be contained in a solid vessel, and thus the expansion of the liquid when heated is complicated by that of the vessel. The expansion actually observed is called the apparent expansion of the liquid. It is less than the real or absolute expansion of the liquid.

Archimedes' Principle. When a body is immersed in a fluid it loses weight equal to that of the fluid displaced.

Area. The extent of the surface of a body.

Barometer. An instrument for measuring the pressure of the atmosphere.

Boiling Point. See *Ebullition*.

Boiling Point of Water. Salts dissolved in water raise the boiling point.

Boyle's Law. The temperature being constant, the volume of a given mass of gas varies inversely as the pressure to which it is subjected.

Caloric. According to the old and abandoned theory, heat was considered to be a weightless, invisible, elastic fluid, called caloric, which was absorbed by bodies during a rise of temperature, and given out during a fall of temperature.

Change of Volume on Melting. Paraffin wax, gold, silver, and many other substances expand on melting and contract on solidifying. Ice, cast-iron, brass, and several other substances contract on melting, and the liquids formed expand on solidification. (Ice contracts on cooling.)

CHARLES' LAW (GAY-LUSSAC'S or DALTON'S). The volume of any given mass of gas kept at a constant pressure increases for each rise of temperature of 1° C. by $\frac{1}{273}$ of its volume at 0° C.

COHESION. The attractive force which holds together the particles of a body.

COMPENSATED PENDULUM. If the pendulum of a clock is made of one metal only, the expansion of the rod with a rise of temperature causes it to swing more slowly, and thus alters the rate of the clock. To avoid this, in a compensated pendulum the length is maintained constant by a suitable combination of metals.

COMPRESSIBILITY. The property in virtue of which the volume of a body can be made smaller by pressure.

CONDENSATION. The reverse process of evaporation; it is the return of vapour to the liquid state, and is brought about either by a fall in temperature or an increase of pressure. Condensation of atmospheric water vapour produces cloud, rain, mist, dew, snow, etc.

CONDUCTION (*leading*). The process by which solids are heated by the vibrations of their particles.

CONVECTION (*carrying*). The process by which fluids are heated by means of moving currents of more or less dense matter.

DELIQUESCENT (*becoming moist*). Substances which absorb moisture from the air, and dissolve in it ultimately.

DENSITY. Mass per unit volume of a substance.

DENSITY, RELATIVE. See *Specific Gravity.*

DEW. Dew is moisture condensed as small drops from the atmosphere on the surfaces of cold bodies.

DEW-POINT. The temperature at which the air is saturated with the water vapour it contains. The temperature of the air at which dew begins to be deposited.

DISTILLATION (*to trickle down*). The volatilisation and subsequent condensation of a liquid.

DUCTILITY (*being drawn out*). The capacity of a solid for being drawn out into wire.

EBULLITION OR BOILING. The rapid conversion of a liquid into vapour, the vapour being formed not at the surface, but within the liquid itself. The temperature at which ebullition occurs is called the boiling-point of the liquid.

ELASTICITY (*extensibility*). The property by virtue of which a body, on being deformed, tends to regain its original shape and volume.

EVAPORATION. The process by which a liquid gradually changes into vapour *at its free surface.*

EXPANSION. Almost all bodies, whether solids, liquids, or gases, increase in size on being heated. In the case of solids we have—

(*a*) linear expansion or expansion in length,

(*b*) superficial expansion or increase in surface area,

and (*c*) cubical expansion or expansion in volume,

while we deal only with cubical expansion in the case of fluids.

FIXED POINTS OF A MERCURY THERMOMETER. Two points of reference, corresponding to two arbitrary temperatures, which are adopted in order to form a scale of temperature. They are:—

(*a*) the temperature of melting ice,

(*b*) the temperature of the steam above water boiling under standard atmospheric pressure.

FORCE. That which changes, or tends to change, the state of rest or uniform motion of a body.

FREEZING POINT OF WATER. Salts present in solution lower the freezing point of water.

FUSION OR MELTING. The change from the solid state to the liquid state. The temperature at which it takes place is called the *melting point* or *fusing point* of the substance. The reverse process is called *solidification.*

GRAM. The mass of 1 c.c. of pure water at a temperature of 4° C.

GRAVITATION (*weight*). The force of attraction operating between masses.

HARDNESS. The resistance of the particles of a solid to any change in their relative positions, such as would be produced were the body scratched.

HEAT. A form of energy; it is the energy possessed by the vibrating molecules of a body.

INFLUENCE OF PRESSURE ON BOILING POINT. The boiling point of a liquid is always raised by an increase of pressure.

INFLUENCE OF PRESSURE ON MELTING POINT. A substance which contracts on melting, like ice, has its melting point lowered by an increase of pressure. A substance which expands on melting, like paraffin wax, has its melting point raised by pressure.

LATENT HEAT (*hidden*). The number of units of heat required to *change the state* of 1 gram of a substance, no temperature change taking place.

LATENT HEAT OF FUSION OF ICE. The number of units of heat absorbed by 1 gm. of ice at 0° C. in changing into water at 0° C.

LATENT HEAT OF VAPORISATION OF WATER. The number of heat units absorbed by 1 gm. of water at 100° C. in changing into steam at 100° C.

LAW OF FLUID PRESSURE. Fluids transmit pressure equally in all directions.

LITRE. 1 cubic decimetre, 1000 c.c. The volume of 1 kilogram of pure water at 4° C.

MALLEABILITY (*property of being hammered*). The property by virtue of which solids may be hammered out into thin plates.

MASS. The amount of matter contained in a body.

MASS, BRITISH UNIT. The pound, a piece of platinum at Westminster.

MASS, METRIC UNIT OF. The kilogram, a piece of platinum at Paris.

MATTER. That which affects our senses, possessing *weight*, *inertia*, and occupying *space*.

MAXIMUM DENSITY OF WATER. The temperature of maximum density of water is 4° C.

MELTING POINT or FUSING POINT. See *Fusion*.

METRE. The metric unit of length. It was originally intended to be the 10,000,000th part of the distance from the North Pole to the Equator, measured along the Meridian of Paris. See page 2.

OPISOMETER. An instrument used for measuring the length of a curved line.

PRESSURE OF ONE ATMOSPHERE. About 14·7 lbs.-wt. per sq. in.

REGELATION (*freezing again*). Ice liquefies under pressure and the resulting water spreads over the surfaces pressed together. When the pressure is removed the water freezes again and unites the pieces.

SOLIDIFICATION. See *Fusion*.

SPECIFIC GRAVITY. The ratio of the weight of any volume of a substance to the weight of an equal volume of water. Density relative to that of water.

STATES OF MATTER. Matter can exist in three states: (1) the solid state, (2) the liquid state, (3) the gaseous state. Many forms of matter may exist in all three states under ordinary conditions.

SUBLIME (*to elevate*). To cause a solid to pass direct to the vaporous condition. On condensation a *sublimate* is formed.

TEMPERATURE. That condition of a body which determines whether it will communicate heat to, or receive heat from, another body when placed in contact with it.

TENACITY (*holding together*). The property of a solid by virtue of which it resists rupture.

THERMOMETER. An instrument designed to measure temperatures. In the construction of a thermometer, some property of matter, which varies continuously with the temperature, is utilised. Usually the expansion of liquids, such as mercury and alcohol, is made use of in the construction of these instruments.

UNIT OF HEAT. The amount of heat required to raise the temperature of 1 gm. of water through 1° C.

VERNIER. A short scale which can slide freely along the graduated edge of an ordinary scale. It enables us to measure to a greater degree of accuracy than is possible with the ordinary scale alone.

VISCOSITY. The resistance particles of a fluid experience in sliding past one another. Fluid friction.

VOLATILE (*winged*). Easily turned into vapour.

VOLATILISATION. The process of turning a substance into the gaseous state.

VOLUME. Amount of space occupied by a body.

WEIGHT. The weight of a body is the force with which the earth attracts the quantity of matter in the body.

YARD. The British unit of length. It is the distance between two fine scratches on gold plugs let into a bronze bar, when the temperature is 62° F.

$\pi = 3{\cdot}14159\ldots$ Ratio between circumference and diameter of a circle. Approximately $= \frac{22}{7}$.

CHEMISTRY

ACID (*sharp*). A substance which combines with a base to form a salt. Has usually a reddening effect on litmus.

ALKALI (*wood ashes*). A substance which affects red litmus, turning it blue.

AMORPHOUS. Without definite crystalline shape.

ANALYSIS (*loosing asunder*). The process of separating a compound into its elements.

ANHYDRIDE (*without water*). An acid oxide, which, *with* water, forms an acid.

ANHYDROUS (*without water*). Dry. Usually refers to bodies without or lacking water of crystallisation.

BASE (*basic oxide*). Usually a metallic oxide, which can neutralise an acid forming a salt.

CALX (*lime*). The earthy residue left when metals are calcined or heated in air.

Combustion (*burning*). Vigorous chemical action, accompanied by the evolution of light and heat.

Compound. A substance formed by the chemical union of two or more elements.

Distillation (*to trickle down*). The volatilisation and subsequent condensation of a liquid.

Efflorescence (*blossoming*). The loss of water of crystallisation when a substance is exposed to the air, involving loss of crystalline form.

Electrolysis. Analysis by electricity.

Element (*a first principle*). A substance which up to now, and by the means at our disposal, has not been decomposed into anything simpler.

Eudiometer (*fine weather measurer*). A graduated glass tube used for determination of volume of gases (originally for the analysis of air, the composition of which was supposed to vary with the weather).

Hydrogen (*water former*). Name given to inflammable air, because it is contained in water.

Incandescence (*being very hot*). The evolution of light by strongly heated particles.

Neutralisation (*neither*). The process by which an acid or an alkali is rendered without effect on, or *neutral* to, litmus—a *salt* being formed.

Nitrogen (*nitre former*). Modern name for inactive air or azote—given because it is contained in nitre.

Oxidation. The process of chemically combining oxygen with a substance.

Oxygen (*acid former*). Name proposed by Lavoisier for active air.

Phlogiston (*burning*). A spirit of fire supposed to escape from combustible bodies when heated.

Reduction (*drawing back*). The process of withdrawing oxygen from a compound containing it.

Solubility. The amount of a substance required to form a saturated solution with 100 grams of the solvent at a particular temperature.

Solution (*loosening*). A uniform mixture of a solid (liquid or gas) with a liquid which is called the *solvent.*

Synthesis (*putting together*). The process of building up a compound from its elements.

Vitriol (*glassy*). Name given to various bodies crystallising in large glassy masses, and hence to oil of vitriol, the parent substance.

ANSWERS

PART I. (PHYSICS)

MEASUREMENT AND MATTER

EXAMPLES I. (Page 6)

1. (1) ·001 metre, (2) ·000001 Km.
2. 8·65 dm.
3. 6750 mm.
4. ·001 Dm.
5. 100,000 mm.
6. 20.
7. 6·436 Km. per hour.
8. 1,093 yds. 1 ft. 10 ins.
9. 11·94 cm.
10. 197·83 ft.
11. 29·22 ins.
12. (1) ·0001 Dm., (2) ·00001 Hm.

EXAMPLES II. (Page 20)

1. 1,000,000 sq. mm.
2. 1,000,000 sq. metres.
3. 176·8 cm., 31,258·24 sq. cm.
4. ·13 cm.
5. 93.528 sq. cm.
6. The areas are the same.
7. 3·545 metres, ·5642 metre.
8. 6·36 cm.

EXAMPLES III. (Page 24)

1. 1,000,000 c.c.
2. 15·625 c.c.
3. 143,600 c.c.
4. 600 sq. cm., 1,000 c.c., 17·32 cm.
5. 804·825 cu. metres
6. 54 litres.
7. 48·125 c.c.
8. 2 : 3.

EXAMPLES IV. (Page 40)

1. 28·416 K.gm.
2. 11·4 gm. per c.c.
3. 8 : 5.
4. 12·5 ins.   5. 10·56 lbs.

EXAMPLES V. (Page 49)

1. 7·35 c.c.
2. The one of larger bulk.
3. 21 gm.
4. 4·375.
5. 2·15.
6. (1) 21 c.c., (2) 2·6, (3) 1.200.
7. ·5.
8. Nothing.
9. 2·58.
10. 22·8 gm.
11. 7·2 inches.
12. ·6 gm. per c.c., 45 c.c.
13. 85 gm.
14. $s_1 : s_2 :: 10 : 13$.
15. 12 lbs. in water, 6 lbs. in the liquid.
16. 2·5.

EXAMPLES VI. (Page 55)

1. 14·51 lbs. per sq. in.
2. 0·91.
3. 33·77 ft.
4. 38·7 ft.

PART II (PHYSICS)

HEAT

EXAMPLES VIII. (Page 72)

1. −40° C., −17·8° C., 82·2° C.
2. 39·2° F., 104° F., 14° F.
3. 167° F., 75° C.

EXAMPLES X. (Page 78)

1. 973·6 c.c.
2. 46·37 c.c.
3. 1085·34 c.c.

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